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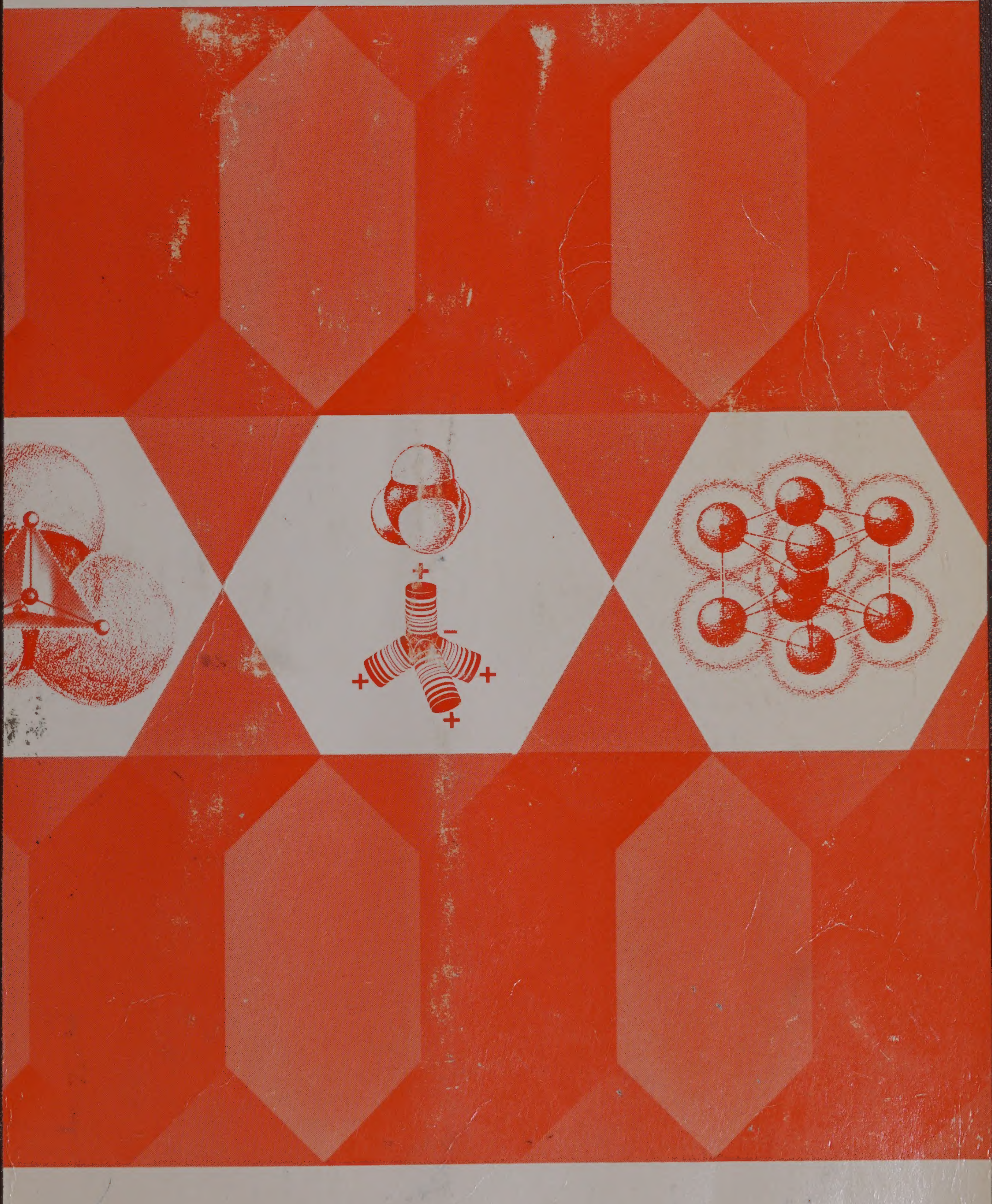
Teachers' Guide

YS TO CHEMISTRY

ELAINE W. LEDBETTER • JAY A. YOUNG



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Teachers' Guide

KEYS TO CHEMISTRY

ELAINE W. LEDBETTER • JAY A. YOUNG

ADDISON-WESLEY PUBLISHING COMPANY
Menlo Park, California • Reading, Massachusetts • London • Don Mills, Ontario

COMPONENTS OF THE KEYS TO CHEMISTRY PROGRAM

Text

Laboratory Manual

Teachers' Guide

Tests

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FOREWORD

The authors of KEYS TO CHEMISTRY have produced a deliberately simple and selective chemistry program. It is designed to provide an understanding of the major principles and concepts of chemistry without losing the student in details.

We believe that this program will allow chemistry teachers to recruit into the study of chemistry many students who would be frightened away by more rigorous, more detailed programs. The authors have succeeded in creating a sequenced laboratory and text that communicates to students the objectives of the course and the behaviors they are expected to learn.

Unlike most conventional chemistry programs and because this program allows individualized learning, both with respect to pace and the pursuit of interests, many of the items that are normally found in a Teachers' Guide have been placed in the text. Chapter objectives, the suggested order of study and a self-test (along with answers at the back of the text) are examples. Placement of these items in the text, along with suggestions for creative work and writing and practice exercises, allows the student to take over management of his (or her) personal study program. He can move ahead on his own with minimum dependence on teacher direction.

A number of other characteristics of the text need to be brought to your attention so that you can use them to full advantage. One of these is the view that students should use other sources of information, in addition to the text. Exercise 1-1 in the Lab Manual is specifically designed to emphasize the existence of handbooks of chemistry. Another handbook which is too often overlooked by all of us is the dictionary. Rather than supply a glossary with the text, we make the suggestion that you provide a dictionary in your classroom and that you encourage your students to use it.

Another aspect of the KEYS TO CHEMISTRY text is the deliberate effort to help students associate chemistry and chemicals with common ordinary things that they encounter every day. For example, we photographed such items as rocks, foods, soaps, detergents, cleansers, and stock bottles of chemicals together for several of the chapter openers. There are photographs of the diverse situations in which the solvent action of water is evident. We also inserted photographs of labels from commonly encountered store items with the hope that students will become curious about the chemical nature of many of the things they use and eat. You can follow up this emphasis by suggesting that students do creative research with the actual substances. You might even permit a party day from time to time if the serving of a cake, cookies, or homemade ice cream was accompanied by an explanation of the chemical and physical aspects of the preparation of such goodies.

Laboratory learning is one of the most important aspects of science. To study science from a book alone provides the student with a distorted concept of chemistry as a science. He will perceive it only as an organized body of knowledge. He also needs to perceive it as a form of learning in which natural things are observed directly so that he can generate knowledge from that observation and also more completely understand the knowledge recorded in the text and elsewhere. As a teacher implementing this course, you need to be very conscious of the learning style that can be built around laboratory work. You will need to be ready to block the student who will try to bypass experience with this learning style by simply reading, and possibly dry-labbing his experiments.

Because the authors believe that laboratory learning is very important, this Teachers' Guide focuses on helping you implement the laboratory program of the KEYS TO CHEMISTRY course. We believe that effective implementation of the laboratory experiences will help the conceptual, theoretical learnings to occur more naturally.

We suggest that you assign students the reading of the preface to the text ("To the Student," p. iii) and the section "About the Authors" (pp. iv, v). The authors have tried to write the text in an informal style, and we have placed an emphasis on them as real people by giving the students background information about them. We hope that you can use these elements to advantage as you introduce your students to the course.

The Editors

Acknowledgements

The Evaluation Cards that accompany KEYS TO CHEMISTRY represent a unique, creative approach to testing. Elaine Ledbetter encountered the idea for the cards at Thornlea School in Toronto, Canada, where their basic character was described to her by Jane Garden.

As KEYS TO CHEMISTRY is used, we hope that other teachers will generate ideas which they will share with the authors for possible incorporation into this program.

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OVERVIEW OF THE COURSE

THE PHILOSOPHY OF KEYS TO CHEMISTRY

Enrollment in chemistry courses has been steadily declining over the past several years. One reason for this decline in enrollment is that textbooks have concentrated heavily on the theoretical and mathematical aspects of chemistry.

Many young people are not science oriented and most do not know what vocation they want to pursue. After leaving school many will have no further opportunity to study science in a formal, organized way; others will bypass opportunities for further study of science. Both kinds of student should be encouraged to explore a wide variety of fields, inside and outside science, as part of finding a vocation that will be suited to their interests and abilities. We have written KEYS TO CHEMISTRY specifically to make the study of chemistry a meaningful exploration for such students.

Science and technology have advanced to the point where man can exert a significant control over his environment. It is imperative that young people develop an understanding of and an appreciation for science if they are to know whether our control of the environment is properly used. It is far more important that students leave their science courses with a firm understanding of the methods of science and a comfortable knowledge of the principles of science than with a partially organized accumulation of facts and misunderstood, half-perceived theories. Also, if the courses they take can generate a feeling of excitement about science and a genuine interest in science, then their science education will continue and will be a source of lifelong pleasure.

The pervading philosophy of KEYS TO CHEMISTRY is not to teach facts and theory, but rather to present chemistry as an important and exciting part of daily living. This program seeks to involve the student in the processes of science as he develops an understanding of the concepts and principles fundamental to the field of chemistry. Many of the models used to explain difficult concepts are drawn from familiar experiences. The importance of individual effort, the trial and error of man through the ages, the role of imagination in scientific progress, and the necessity for effective communication are continuing themes of the course.

Laboratory investigations are an integral part of the program. Each experiment is planned to present a single idea so that the student does not lose sight of the reason for the activity. Emphasis is upon discovery, but sufficient background is provided in the introduction to each experiment to make it meaningful. Some experiments pose questions that the student is not expected to answer at that point in his work. As more information related to the question is obtained, these questions are raised again. This is done to emphasize that we are never completely finished with many problems, and that we need to re-examine earlier conclusions as our understanding increases.

Throughout the course a great deal of emphasis is placed upon individual creativity. Each chapter contains invitations to the student to design experiments to investigate questions he himself has raised; to propose hypotheses for problems arising out of routine experiments; to make predictions and to test them; to plan bulletin board displays; to write original poems and articles; and to develop his creative talent to the fullest extent.

COMPONENTS OF THE KEYS TO CHEMISTRY PROGRAM

THE TEXTBOOK

KEYS TO CHEMISTRY is designed to permit each student to move through the course at his or her own rate. Therefore, each chapter begins with a set of performance objectives. These alert the students to what they will be expected to know and do by the time they complete the study of that chapter. Practice exercises and a self-test are provided at the end of each chapter; answers to these are at the back of the text. The tests are keyed to the performance objectives, so students should do very well on the tests if they utilize these aids to the fullest extent.

Each chapter contains suggestions for creative investigations and for creative writing. Students may work from the suggestions given or they may think of many others, but should be encouraged to do some original activities in connection with every chapter.

The appendices contain many tables and other aids that will assist the student throughout his study of chemistry.

The textbook has eleven chapters. Chapter 1 is introductory and is designed to arouse interest in chemistry as well as to familiarize the student with techniques, equipment, apparatus, and safety procedures.

Chapter 2 deals with the states of matter, with elements, mixtures and compounds, and with energy changes that occur as matter is altered.

Chapter 3 introduces the concept of equilibrium as liquids and the solution process are investigated.

Chapter 4 provides an opportunity for students to become familiar with acids, bases, and salts. The use of chemical indicators and buffers is introduced and their relationship to biological systems and daily activities is stressed.

Chapters 5, 6, and 7 are more theoretical. These deal with the historical development of the modern atomic theory, the structure of atoms, and chemical bonding. Since the student has had sufficient laboratory experience by this time, and since he has been introduced to solutions and acid-base reactions, a great deal of laboratory work is included during the study of these more theoretical chapters. This helps to sustain the interest already generated and makes the study of theory more palatable.

Chapter 8 moves into the quantitative relationships so important to the meaningful application of chemical theory. The extent to which each student is expected to master this chapter is left to the discretion of the student and his teacher. Thus, students who have developed a deep interest in chemistry by this point in the course and who may wish to pursue the study further can do so with this material. However, students who are less interested (or able) should not be forced to do formula writing and equation balancing. It is far more important that these students leave the course with a feeling of interest and excitement about chemistry. Clearly they should not develop a dislike because they thought they were required to spend long hours of drill trying to master a technical aspect that they tend to identify as irrelevant and boring.

Chapter 9 introduces the properties of the gaseous state of matter. Again, emphasis here is upon general principles and concepts and their applications to daily living, rather than upon mastery of the gas laws as such.

Chapter 10 deals with the properties of solids and their solutions. Some of this material ties in with what was learned in Chapter 3, some of it with the quantitative aspects of solution reactions, and some with the structure of the solid state.

Chapter 11 is generally considered by students to be the most exciting and rewarding part of the course. It introduces the student to elementary qualitative analysis. Emphasis here is upon tying together all the major principles and concepts of the course in a practical way. Once again, the extent to which each student is required to write equations to represent his observations should be determined by the interest and the ability of each individual.

THE LABORATORY MANUAL

The first section of LABORATORY KEYS TO CHEMISTRY contains pictures of the apparatus and equipment commonly used in elementary chemistry. Immediately following this section is a discussion of the rules of conduct expected in a science laboratory and of the safety precautions necessary for the protection of the workers. This section is followed by a description of laboratory techniques and procedures. Students should become aware of the techniques presented here so that they may refer to this section as needed.

Some teachers like to have students master the techniques very early in the course before proceeding with formal study. Others prefer to introduce new techniques only as they are needed. The method you follow will be determined by your own preference. Each has its advantages.

Each experiment contains some background information, a statement of the purpose for the investigation, the procedure to be followed, and questions to be answered at the conclusion of the experiment. When they are needed, directions are given for tabulating and graphing data. Most of the experiments are accompanied by illustrations showing how equipment should be set up. When it seems appropriate, special safety precautions are stressed.

Since students are expected to keep their own records, no data tables are placed in the laboratory manual. We strongly recommend that each student purchase a stitched notebook in which he can keep a detailed record of his laboratory investigations. Suggest that he leave the first few pages blank so that he can prepare a table of contents. This will make it easier for him to locate the data from each experiment. Prior to going into the laboratory, the date, experiment number and title, and the purpose of the experiment should be placed in the record book. If data are to be collected, appropriate tabular headings should be entered in the record book ahead of time.

Careful attention by the teacher to the prelab preparation of the record book early in the course will eliminate many problems. Require that each student have his record book initialed by you before going into the laboratory. This has several advantages: (1) the student must do sufficient preplanning in order to prepare proper data tables so that he has a clear idea of what he will do in the laboratory and knows why he is performing the experiment, (2) he has a permanent and continuing record of his investigations, and (3) more opportunity for creative thinking is provided.

After each investigation has been completed, the student should submit a formal report of his results. This should be typed or written in ink and should include:

a. the title and number of the experiment
b. the purpose of the experiment
c. observations and/or data in tabular form
d. graphs (when appropriate)
e. tentative conclusions and/or answers to questions asked in the manual

Whether a formal report on every experiment is to be required is a decision each teacher must make.

We recommend that such reports be required on the majority of the investigations for these reasons: (1) students should learn that one major aspect of the scientific process is effective communication of results; (2) writing the formal report helps the student to clarify the meaning of his data; (3) knowing that a formal report is expected is conducive to more careful laboratory work.

In addition to the experiments, there is one exercise included in the laboratory manual. It is designed to familiarize the student with the *Handbook of Chemistry and Physics* very early in the course. Also, Experiments 6-1 through 6-4 are thought experiments that some people would label as exercises. They are important and have been planned to permit the student to construct his own periodic table of the elements. The process used by the student makes this table meaningful to him in a way that can never be achieved if instead it is presented to him as a finished product.

THE TEACHERS' GUIDE

This TEACHERS' GUIDE contains suggestions for implementing the course. For each chapter, we state goals and give teaching hints, including demonstrations where appropriate. We have included a time schedule for Chapter 1 to help you pace the students at the beginning of the course. And for each experiment the following information is given:

a. the purposes of the experiment from a teaching viewpoint
b. materials required per student or per class, as appropriate
c. quantities of materials required per student, or per class
d. instructions for preparing solutions
e. prelab discussion hints
f. sample student data (when appropriate)
g. postlab discussion hints
h. answers to questions and/or sample data tables and graphs

THE TESTING PROGRAM

The testing program for KEYS TO CHEMISTRY consists of a deck of 330 EVALUATION CARDS, plus a teachers' guide with answers and sample forms for the answer sheets and report sheets (on which the student receives from the teacher an analysis of his test results).

There are 30 cards for each of the 11 chapters. The cards are coded, both by color and by capital-letter code, to indicate the type of material being tested:

Color	*Letter Code*	*Type of Material*
white	L	laboratory practical
green	M	mathematical problem and/or formulas or equations
blue	I	interpretation of data
pink	F	factual information
lavender	AU	application and understanding of principles
gold	H	honor question

For each chapter there are five cards of each color, labelled a, b, c, d, and e. The questions on each of the five cards of one color are as parallel as possible. All test questions are tied closely to the performance objectives. The nonlaboratory questions are similar to the practice exercises and self-tests in the text.

To take a test, the student draws *one* card of *each* color (except gold). Thus, a maximum of five students can take a test on the same chapter on the same day. When the course is self-paced, one deck of cards is sufficient for the entire class. The cards are re-usable and should last for the lifetime of the text. A protective coating makes it easy to remove stray marks.

The student taking a test is given a printed answer sheet and a kit for the laboratory question drawn. (The lab kits can be made up in advance in tote boxes, as we suggest for the regular laboratory program.) On the answer sheet the student is asked to list the complete code number for each card drawn. The code number "3-I (d)," for example, is on a blue card (interpretation of data) for Chapter 3; the "d" identifies *which* blue card for Chapter 3. To correct the test, the teacher looks up the answer for that code number.

After correcting the test, the teacher fills out a report sheet which gives the student's score and advises the student whether to proceed to the next chapter or review the chapter and take another test on it. The report sheet for each chapter lists the topics covered in the chapter, so the teacher can check off what topics in particular the student should review.

When repeating a test, students draw from the same deck of cards. If they happen to draw one or more of the same cards, we do not consider this to be a disaster.

If the student makes a certain score on his *first* test attempt, he is eligible to draw an honor (gold) card. The honor questions are more theoretical and pursue topics at a deeper level than the regular questions.

INDIVIDUALIZED INSTRUCTION

For many years group learning has been accepted as the best method for presenting information to students. But as teachers have become more cognizant of the wide range of ability and achievement of students, even in carefully grouped classes, the need to individualize instruction has become evident.

Individualized instruction takes many different forms. Ideally, each student should be able to move through a course at a rate that is comfortable for him. Provision should be made for supplementary work in areas where the student has insufficient background, for enrichment material to challenge the most able student, and for interaction among the various members of a class.

Education becomes a more meaningful process when a student assumes the responsibility for his own learning and becomes involved in the process of planning his own schedule of learning activities and setting his own goals. If he must compete, he can compete only with himself. The promise of immediate advancement is always present, and the fear of failure is removed.

THE ROLE OF THE TEACHER

When the individual progress plan is followed, the role of the teacher becomes quite different from that assumed under group learning. The teacher is now a consultant and an advisor, not a dispenser of information. In order to serve in this new capacity the teacher must not only have a broad knowledge of the subject matter, but must also acquire a deeper insight into the interests, capabilities, and needs of each student.

Teachers who use this plan for the first time face an adjustment period if they are accustomed to playing the dominant role in the classroom. Many teachers enjoy their role as the star of the show and find it difficult to relinquish. The organized confusion that prevails when students are engaged in a variety of activities may be disturbing to teachers who are accustomed to absolute quiet. However, the advantages of this plan are sufficiently great to warrant that each teacher give it a try.

THE ROLE OF THE STUDENT

Students who have never been given the freedom to work at their own pace also require an adjustment period. It may take some quite a long time to make the transition from teacher-directed learning to self-pacing. Some students will need assistance and encouragement during this period. For those who find it difficult to assume the responsibility for moving ahead at a reasonable rate, the teacher may need to help the student plan a daily schedule for himself. As he gains confidence in his ability to progress on his own, such help can gradually be withdrawn.

The shy, insecure, or retiring students require special attention. Teachers need to be particularly alert to their needs, because these students will not otherwise seek the help they must have. They may feel hopelessly abandoned unless the teacher sits down with them and assists them individually in organizing their plans. By providing individual attention to their needs, by speaking an encouraging word at opportune moments, and by an attitude of warm interest, the teacher can help each student build confidence in his or her own abilities. Unless a teacher is aware of these students and seeks them out, the more aggressive and demanding students can easily absorb all the teacher's time without his being conscious of the neglect of the quiet ones.

Students who are accustomed to making high grades by memorizing material and by doing exactly what the teacher wants will need very special attention. Some will feel extremely frustrated and lost when placed in the unfamiliar situation of directing their own learning. They will ask such questions as, "What do you want me to do?" "How much material do you want me to cover this grading period?" and "Just what must I do to make an A?" You will have to help them learn to ask these questions of themselves: "What should I be doing?" "How much material should I cover?" "What do I need to do to learn this material well enough to rate a grade of A?" A friendly discussion concerning their purposes, their future goals, and their freedom to pursue the study of chemistry as rapidly and as deeply as they desire will soon resolve their concerns and put them at ease.

ORGANIZATION OF THE COURSE

It is recommended that classes proceed through Chapter 1 as a group. This enables students to become familiar with laboratory apparatus, equipment, and procedures. Instruction in safety which includes not only procedures but also the location and operation of such equipment as fire extinguishers, shower, eyewash fountains, etc. is most important and can be expeditiously accomplished by the group method of instruction.

During this period of orientation students must be told that they are preparing to move ahead at their own rate and that at the end of Chapter 1 they will assume the responsibility for their own planning and progress. This provides motivation for them to learn how to conduct experiments, write formal reports on laboratory results, and to get a feel for the type of work that can be done only in the classroom. Setting a fairly rapid pace through the first chapter creates interest and provides an impetus which helps carry students through the adjustment period as they begin to work on their own.

Another advantage of the orientation period is that students can be made aware of the importance of the performance objectives and their relationship to the evaluation that will be made at the close of each chapter. They should learn very early in the course that the objectives, the practice exercises, and the self-tests are all directed toward helping them master the principles and concepts in that chapter. Also this is the time at which to point out the many aids in the appendices of the text.

COVERAGE OF MATERIAL

We feel that in order to receive credit for the course a student must complete a minimum of seven chapters. Chapter 11, which deals with elementary qualitative analysis, is required of all students and is taken during the last four weeks of the school term. All students are required to study this chapter at the same time because of the large number of solutions needed for the laboratory work.

Students should be told about Chapter 11 at the beginning of the course. Those who complete the entire course up to Chapter 11 before the time scheduled to begin this study can be given several options. They may be free to study topics of their own choosing for extra credit. If they do not wish to do this, they may be excused from class attendance until the study of Chapter 11 starts. In no case, however, should they be required to do supplementary work, or in any way be made to feel that they will be penalized for early completion of the course. Rather, there should be some tangible incentives for working up to their full capacity.

Minimum work should be set for each grading period. Otherwise, some students will fall so far behind during the early part of the year that it will be impossible for them to complete the course minimum of seven chapters. If class periods are one hour in length and if the class meets daily, then a minimum of 1½ chapters is reasonable for a six-week period. Alternatively, a semester minimum of 4 chapters could be required, but for the students this needs to be broken down into a minimum weekly or biweekly schedule.

Recognition of learning in a course like chemistry has been traditionally expressed in the form of grades and credits. Usually the amount of credit is the same for all students and is associated with a semester or a school year. The variable with which the teacher has been able to work has been the grade.

Under an individualized program, the constancy of credit for all students becomes undesirable. The variation in grade alone becomes a very inadequate and restrictive channel for describing student accomplishment.

If your school and its manner of organization will permit it, you might consider variation in credit as well as in grade for this individualized chemistry course. If ten credits are ordinarily associated with a one-year course, then each chapter could be rated as worthy of one credit at a grade that reflects the quality of the work done on it. If a fast-working, able student completed ten chapters in considerably less than the allotted year of time, he could then be free to use his time on other studies or he could continue working on chemistry, and by doing so, he could earn additional credits in chemistry.

Taking the above approach to grades and credits will resolve some problems. Further, it is an inevitable consequence of individualized instruction.

PROVISION FOR VARIETY

As soon as students begin to move at their own rates, some grouping appears. Pairs may form of students who like to work together. This is especially evident in laboratory activities. Larger groups of from 3 to 7 students may form for discussion of textbook material. We have found that these groups, especially those formed spontaneously for discussion purposes, tend to be very flexible. They continually change from day to day and from week to week. It is not uncommon for the more able students to lead these discussions. The teacher may or may not be called into a group for consultation. If he is not needed, then he should not impose his presence upon the group.

Variety in routine is extremely important in order to maintain high motivation. There are several ways in which such variety can be introduced into a classroom using self-pacing.

a. Special Presentations

The class can be convened as a group at intervals for a special presentation. All such interruptions of the routine should be scheduled well in advance so that students can account for them in their individual planning. These presentations can be a special demonstration, the showing of a film, a guest speaker, or a lecture by the teacher.

b. Problems

Another variation is to set up a problem not necessarily related to any unit of study, but which has high interest value. Some examples are:

1. The Drinking Duck

This toy can be purchased for one or two dollars in most hobby stores. Set him up and let him be drinking as the class assembles. Pose such questions as:

Why does Albert drink? (It lends interest to give him a name.)

What factors might affect his drinking rate?

What scientific principles are involved in his activities?

Announce that he will be available for observation for a few days. During this time, students may only look, not touch. Encourage them to submit papers proposing hypotheses and experimentation to answer the questions stated above. Perhaps a time limit can be set for submission of these papers. You might offer a "Nobel Prize" for the best papers.

Then, after submission and discussion of the papers, permit students to investigate their hypotheses. This can be done as a class activity so that all may observe the results and enter into the discussion.

2. Dancing Moth Balls

Obtain a tall transparent cylinder, such as a 250 cc graduated cylinder. Fill it with water. Add about ½ teaspoon of citric acid and about 1 teaspoon of bicarbonate of soda. Shake well. Drop 6 to 8 moth balls into the water. Have this set up and working prior to class. As students come into the room, they will be attracted by the fact that the moth balls are moving up and down in the water. This raises many questions:

Why do the moth balls move up and down?

Is water the liquid in the cylinder?

What causes the bubbles on the moth balls?

Again, permit observation only for a few days. Encourage the submission of hypotheses and suggestions for testing them. After an appropriate period of time has elapsed, permit individual experimentation in the laboratory. Before individual experimentation has begun, require that each student who wishes to experiment submit his proposal. This should include his purpose, materials needed, and step-by-step procedure. Check his proposal for safety only. Then let him proceed.

c. Communication and Sharing of Experiences

We feel that it is important that students working individually have frequent opportunities to come together as a group to communicate with each other and to share ideas. This can be accomplished in several ways. Two are:

1. Current-Event Discussion

Many exciting things are happening in the world which involve science and technology. Set aside a day every two or three weeks on which students can report on such an event. This provides an opportunity to include the societal implications of such happenings. Economic, political, and ethical considerations may also be pursued as is appropriate.

2. Discussion of Student Investigations

Since much original laboratory investigation is encouraged in this course, frequent opportunity should be given for students to report on such investigations for the benefit of the class. This may take the form of oral presentation accompanied by actual demonstration of results, techniques, etc., and while it is especially helpful for the shy, retiring student, overly aggressive students can learn more about how to get along with their peers.

d. Films

Film showing presents a problem when students are moving at their own rates, because they become fairly diverse in their interests as the course develops. Unless one has a school library of chemistry films which can be shown as needed, it is difficult to schedule them at a time that is equally appropriate for all students. However, films illustrating key concepts should be scheduled at frequent intervals during the term. Admittedly, some students may be ahead of the concept and others will not have studied it at the time, but some attempt must be made to weave good films into the course.

Many films of general interest are available which can stimulate interest in science and technology and can be used at any time. These also serve to provide variety and maintain interest.

ORGANIZATION OF THE LABORATORY

Good organization and advance planning is essential for smooth operation of the laboratory when the individual progress plan is used. Each student should have a locker in which all commonly used equipment is kept. The reagents and special materials for each individual experiment can be placed in tote boxes each clearly marked with the number of that experiment. These tote boxes can be kept on a central supply table so they are easily accessible to students when they need them. When a student is ready to perform a given experiment, he goes to the supply table, picks up the appropriate tote box, takes it to his laboratory station, and uses it. Upon completion of the experiment the student returns the box to its proper place on the supply table.

Once the tote boxes have been prepared for each experiment, little time is required to keep them replenished. It is assumed that students will take responsibility for keeping their own laboratory stations neat and clean.

Early in the term it may be necessary to have several tote boxes for each experiment, but as students become spread out in their work, one box per experiment is usually sufficient.

It is very desirable to have someone in the laboratory at all times to check on safety procedures and to supply needed equipment. If paraprofessional help is available, this is the ideal arrangement. However, if such help is not available, students may be used. If you can enlist the aid of a student who has had chemistry and who is responsible and knowledgeable, he can be a great help in keeping the laboratory running efficiently. Sometimes credit can be given for such assistance and is an added incentive for students to work with the teacher.

CREATIVE WORK

KEYS TO CHEMISTRY places much emphasis upon creative work by each student, particularly upon student-designed labs. The student designs and carries out these labs after his plans have been approved by the instructor. Upon submission of a formal report on his results he receives credit for the work.

Student-designed labs may arise from questions raised by the required experiments, from class discussion, from outside reading, or from questions posed by the student himself. In order to conduct this kind of lab, students should submit a plan at least 24 hours prior to the period when the experiment is to be done. The plan should include:

a. purpose (or hypothesis)
b. materials and/or equipment required
c. step-by-step procedure that he intends to follow
d. safety precautions where appropriate

The teacher needs to check the plans for safety. If the plans are acceptable from the standpoint of safety, they are usually approved. Of course, in some cases suggestions may be made by the instructor, but in general we suggest that the student be permitted to discover weaknesses in his plans by experience. When a student does need to be informed of weaknesses in his plan, skillful questioning about procedures is recommended, rather than direct telling. Raise questions in the mind of the student; then let him decide whether or not to modify his original plans.

If students are to capture the spirit of science, they must learn how to ask appropriate questions, how to devise good experiments, and how to evaluate their results.

Other forms of creative work may include posters, art work, and bulletin board displays appropriate to the topic being studied. Creative writing in the form of articles, essays, poetry, crossword puzzles, etc. can also be encouraged. Creating cartoons and writing science fiction are other avenues of expression that students may wish to explore.

Reports on topics of interest to the student obtained from current journals and reference books are not considered creative work unless the student adds his own creative ideas. Such reports can be accepted for extra credit as outside reading.

The encouragement of creative endeavors is too frequently neglected in science courses. All of man's scientific advancement has depended upon originality of thought and expression. There is no better place to stimulate the development of creativity than in science courses. Indeed, many students will need a certain amount of prodding before they will undertake creative work, but once started they not only derive enjoyment from these endeavors, but frequently find that they have real talent. This self-discovery is a viable and fruitful product of this course.

EVALUATION

Evaluation should include objectives from both the affective and the cognitive domains. Factors in the affective domain are necessarily scored subjectively, but it is especially important that these be included in the evaluation of students in an individualized course. Some of the factors to be considered are: (1) the attitude of the student, (2) interest shown in his work, (3) open-mindedness, (4) willingness to accept constructive criticism, (5) resourcefulness, (6) respect for others, (7) care of equipment, (8) cooperation, (9) willingness to assist others, and (10) use of time.

It is recommended that these factors be discussed with students in the opening days of the course. The students should each have a copy of these objectives so that they can evaluate themselves.

Testing

When the individual progress plan is used, the EVALUATION CARDS (described earlier) are especially convenient and appropriate. A student may take a test on any day that he wishes. It is suggested that all students wishing to take a test during a given period present themselves as soon as roll is checked. This allows them to use the entire period if necessary, and also frees the teacher to work with others in the laboratory or in discussion as needed.

To take a test the student draws *one* card of *each* color (except the gold). He is given a printed answer sheet and the kit for the laboratory question, and proceeds to take the test without consultation with other students. The time for completing a test varies, but most students usually complete the tests within 30 minutes. However, no other time limit should be imposed except that the test must be finished within the class period.

Scoring and Grading

The maximum score on any test can be set at 10 points, with each card counting 2 points. Partial credit may be given on any card if the answers are partly correct. Under this scoring system, a student must score a minimum of 5 points in order to proceed to the next chapter. No student is permitted to begin a new chapter until he has passed the test on the preceding one. Each student is given a report sheet telling him his score and advising him to proceed to the next chapter or to review the topics checked on the sheet.

If the student fails to make 5 points on the test, he reviews the material on which he was weak and takes another test (drawing from the same set of cards) when he feels he has mastered the topics. Some students will request an immediate retest. This request should, under most circumstances, be denied. We feel that if a student cannot score 5 points on a test, then some time is needed for sufficient remedial work to make another test attempt worthwhile. An exception might be made if a student is repeating a test and scores 4 points and it is at the end of a grading period. Sometimes, a very weak student who is trying to the full extent of his ability may be given a 5 if he has failed to pass after having tried two or three times. This is especially true on Chapter 8 with students who are not oriented to a further study of chemistry. The teacher must use his best judgment in these cases, considering the attitude, ability, and effort of each individual student.

Two devices can be used to encourage students to do their best on the first test attempt: (1) If a test must be repeated, the maximum score that can be recorded is 5 points. (2) Only if 8 or more points are scored on the *first* try can the student draw an honor card.

If a student makes 8 or more points on his first test attempt, he becomes eligible to draw an honor (gold) card. Honor cards may count as follows: (1) If the card points are to be converted to number grades, as indicated in the next paragraph, then an honor card counts a maximum of 10 points, or enough points to bring the number grade up to 100. (2) If the point system is used (see Methods for Determining Grades), then the honor card may count from 0 to 10 points on the grading period total. The student must understand that if he draws an honor question, he can never lower his original grade on the test. Students may use any references they wish in answering the honor questions, but they must work without discussion and must complete the question within the class period.

For grading purposes the following scale can be used:

Points on test	*Number grade*
5	70
6	75
7	80
8	85
9	90
10	95

When honor points are added to a test score, no grade above 100 is given.

Methods for Determining Grades

As a rule of thumb, a student should cover an average of 1½ chapters each 6 weeks. Students aiming to complete the course should aim for 2 chapters every 6 weeks. For credit in the course, a minimum of 7 chapters should be completed. (Non-science-oriented students may omit Chapters 6, 7, 8, and 9.) There is no completely satisfactory system for determining the grade a student earns in a grading period. However, three different plans that we and others have used with KEYS TO CHEMISTRY during the pilot period are summarized here.

Plan I: A Point System

In this plan the number of points accumulated during the grading period determines the grade a student receives. As an example, points can be distributed in this manner:

Formal reports on laboratory work: 2 to 5 points each
Outside reading reports: 2 points each
Chapter tests: 5 to 10 points each
Affective domain: 10 points maximum
Creative work:
  Free labs: 10 to 15 points each
  Bulletin board display: 5 to 10 points each
  Posters, cartoons, puzzles, etc.: 3 points each
  Essays, articles (minimum of 2 pages): 8 points each

Suggested point values for formal laboratory reports using this system are:

A = 5 points
B = 4 points
C = 3 points
D = 2 points

Note: We have found it most useful to give an "incomplete" on reports that are poor, rather than to give a C or D. When the report is incomplete, the student must do more work on it and this results in a better understanding of the experiment. If a C or D is given, the report is usually filed and no further study is done. In case of an incomplete grade, the maximum grade that can eventually be earned is B.

Example: Maximum number of points obtained after completion of Chapter 1. (Choices of creative work subject to considerable variation.)

5 laboratory reports (5 points each)	25
3 outside reading reports (2 points each)	6
test	10
affective domain	10
subtotal	51
creative work:	
3 free labs (15 points each)	45
1 bulletin board display (10 points)	10
3 posters (3 points each)	9
1 poem (3 points)	3
1 essay, article, etc. (8 points)	8
subtotal	75
total	126

The following table is a suggested basis for converting points to letter grades for a 6-week grading period.

Grade	*Total points earned in 6 weeks (minimum)*	*Minimum points from creative work*
A	115	50
B	90	30
C	65	15
D	50	10

Plan II: The System of Thirds
In this plan, reports on experiments, test scores, and other work each count one-third of the final grade. Other work includes free labs, outside reading reports, and creative work.

Plan III: Student Self-evaluation
No one is in a better position to know how hard a student has worked, how much creative work he has done, and what he has actually accomplished than the student himself. We have found it most rewarding to provide students with a check sheet at the end of the grading period and ask him to assign the grade that he feels he has earned.

The check sheet has a place for the student to list the experiments he has performed and the grade he has earned on each; the number of tests taken and the score on each; and the creative work done and the value he places upon it. In addition, a place is provided for the student to write a brief evaluation of the total work he has done during the grading period.

One mark of an educated person is the ability to evaluate objectively his own strengths and weaknesses without over- or underestimating them. This requires experience. What better place to gain such experience than in a course in which you assume the responsibility for your own learning?

PROBLEMS AND POSSIBLE SOLUTIONS

In any course, there will be problems. This is true regardless of whether one teaches by the traditional method or whether one uses the individualized approach. We will try to smooth the way for you by mentioning a few recurring problems along with possible solutions.

1. Unwise Use of Time by Students

There are certain days for each of us when we do not feel like working. Students are no different. Therefore, it can be expected that at times there will be students who accomplish very little during a class period. An occasional lapse of this kind is no cause for alarm. A kind remark accompanied by a smile that indicates understanding on the part of the teacher is sufficient to take care of these cases.

Prolonged waste of time is a different matter. In some instances, perhaps groups need to be broken up for mutual benefit. In other cases, it may be that the student needs help in establishing a realistic daily schedule. If this is true, then the teacher should sit down with him and provide assistance in writing such a schedule and urging him to follow it until better work habits are formed. For others, a personal conference will help to solve the problem. This is particularly true if the conference is conducted in a friendly, non-threatening atmosphere. If the student feels that he can trust the teacher and that the teacher is genuinely interested in helping, most students will respond. Everyone wants to do well, and all of us need assistance over our personal rough spots.

If these efforts fail to produce effective use of time in class, then sterner methods may have to be employed. For example, requiring a student to spend an hour before or after school doing what should have been done in class can be effective.

2. Failure to Work up to Capacity

It is not unusual to find capable students who are doing but a fraction of what they could be doing. Through casual conversation find out what interests them; find out if they have personal problems; be open and friendly with them. If you can gain the confidence of such students, then the path is open for an honest exchange of opinions. Many times if a teacher is frank and lets the student know that his teacher is aware of the fact that he is operating at a low level of achievement, they can work out the problem together. Perhaps the student is not being sufficiently challenged; perhaps he is bored; or perhaps he has discovered that he just is not interested in chemistry. Whatever the reason, an effort to arouse his curiosity or otherwise stimulate him is appropriate.

For example, sometimes good students can be enlisted to help less able ones as a means of encouraging the good student to work to his fullest capacity. Asking a student who could do better to aid you in dispensing lab materials, keeping the room in order, or helping to prepare kits for experiments will often stimulate him.

3. Dry Labs

Occasionally one finds a student who will write lab reports without having performed the experiment. The best way to combat this practice is to require that each student have his data table in his lab record book initialed by the instructor prior to going into the lab and again upon completion of the experiment. Final initialing should be done while the student can exhibit his results. This also provides an opportunity to confer with each student about his observations before he discards his reagents and solutions. Often errors and misunderstanding can be caught in this way early enough to enable the student to correct them prior to writing the lab report.

4. Ghost-Written Creative Work

We have not found this to be a problem, but it could become one. Gaining the cooperation of students, establishing good rapport with them, and letting them know that you believe that they can produce original and creative materials will go a long way towards eliminating this possibility.

5. Simultaneous Demands on the Teacher

If one teacher must supervise a class of 24 or more students, there are bound to be times when he is needed in the laboratory, by a small discussion group, and by a student taking a test all at once. At other times several discussion groups may require his assistance at the same moment. This can be a real problem, frustrating to both students and teacher. When such occasions arise, explain that you will take each in his proper turn and that you are sorry that some must wait a few minutes for help. Often another student may be able and willing to assist. Every effort must be made by teachers to secure paraprofessional help in laboratory classes. Administrators must become more aware of this legitimate and growing need. The use of volunteers who are willing to help without compensation is becoming quite common in some communities.

ADVANTAGES OF THE INDIVIDUALIZED APPROACH

There are many advantages and rewards for both students and teachers who use the individualized approach to learning. A partial list of these advantages are:

1. Students grow in their ability to direct their own progress in education. When this happens they are far more likely to keep on learning throughout their lives.
2. Teachers and students come to know each other better than in the more formal classroom. They work together on a one-to-one basis that is conducive to mutual understanding.
3. Most students complete more work and go into greater depth than they do in a conventional classroom.
4. Fear of tests is eliminated. A student knows what he will be expected to do and can prepare for the test. He knows that experiments, textbook material, practice exercises, and the self-test are all designed to help him to achieve the stated objectives. He competes only with himself.
5. The slowest and least interested students accomplish *something* and achieve a certain sense of satisfaction. Even those who do not meet the minimum requirements for credit know where they stand at all times.
6. Real failures are reduced to nearly zero. Generally, students who do drop out do so because they do not want to assume the responsibility necessary to discipline themselves. All who persist and meet the minimum requirements make a passing grade with very, very rare exceptions.
7. Students learn how to organize their time, how to pace themselves at a rate that is comfortable for them, and how to evaluate their strengths and weaknesses.
8. Students learn independence. They become familiar with reference materials. They learn to propose hypotheses to answer their own questions and how to design experiments to test these hypotheses.
9. Because they constantly use the methods of science themselves, students acquire an understanding of and an appreciation for science.
10. Teachers have an opportunity to work with students individually. This provides a unique opportunity for drawing the student out, leading him to think more deeply, and really getting to know him as a person. The teacher will develop a genuine interest in each student which will lead the student to the realization that the teacher cares about him as a person.

CHAPTER NOTES

CHAPTER 1

THE A, B, C's OF CHEMISTRY: ATTITUDES, BUILDING BLOCKS, AND CONCEPTS

Goals:

1. To emphasize the attitudes and procedures of scientific investigation.
2. To develop familiarity with laboratory apparatus and skill in using it.
3. To introduce routine laboratory techniques and safety procedures.
4. To develop skill in keeping good records on laboratory investigations.
5. To develop skill in measurement and to emphasize the importance of recording the degree of uncertainty in measurement.
6. To introduce graphing techniques.
7. To introduce the concepts of density, equilibrium, and randomness.
8. To prepare students to work independently and to move at their own rate.

Procedure:

First Day. As soon as preliminary enrollment duties are cared for, issue textbooks and laboratory manuals. Let students go into the laboratory and learn the names of the various pieces of apparatus by comparing them with the pictures in the laboratory manual. Discuss the kind of laboratory record book each student should purchase. Explain that it should be a stitched, permanently bound book so that a continuing record can be kept of all experimental work. Ask students to bring the book to class on the second day.

Assign for the following day: The study of "Keys to the Chemistry Laboratory" in the laboratory manual (pp. 1 through 4). Have students write the *reason* for each of the conduct and safety rules in their lab record book.

Second Day. Discuss the reasons for the conduct and safety rules. Demonstrate the operation of the gas burner.

Assign for the following day: The study of "The Gas Burner" in the laboratory manual (pp. 5 to 7). Prepare lab record book to enter observations on the operation of the gas burner and related activities.

Third Day. Students will go into the laboratory and follow the procedure in the laboratory manual for the operation of the gas burner and do the related activities.

Assign for the following day: The study of Sections 1-1 and 1-2 in textbook (pp. 2 to 9).

Fourth Day. Discuss the assigned sections. Encourage students to provide their own examples of observations, inferences, hypotheses, etc.

Demonstration – The Blue Bottle

Purpose: To arouse interest in chemistry by showing the importance of making careful observations, proposing hypotheses, and planning tests to check these hypotheses.

Preparation: Prior to class, prepare a blue bottle in the following manner. Secure a 500 cc Erlenmeyer flask. Add about 250 cc of water. To the water add 5 grams of dextrose and 5 grams of sodium hydroxide. Shake well to dissolve the solids. Then add 1 cc of methylene blue solution. (To prepare solution, add 1.48 g solid dye to 100 cc of 95% ethyl alcohol. Let stand for two days, filter, and store for use.) Shake and seal the flask tightly with a stopper.

Activities: As class begins, have the bottle sitting on the desk. The contents are colorless. Lift it carefully and ask students to tell you its color. If it is observed closely, a blue film can be noted at the surface of the liquid, but this may not be evident to students in the classroom.

Shake the bottle. The solution turns blue. Allow it to sit undisturbed and the blue color disappears. Ask students to propose hypotheses to explain their observations. Does the length of time of shaking increase the length of time the contents remain blue? Does vigorous shaking produce a deeper blue than gentle shaking? Will a gentle swirling cause the contents to turn blue? What will happen if the flask is left until tomorrow?

Temporary Conclusions: Do not attempt to answer these questions, but encourage students to do so. This is an excellent way to arouse curiosity and stimulate interest.

It is suggested that you leave the bottle until the following day. Prepare a second, fresh bottle for the following day for comparison purposes. How far you carry this activity and the way you conclude it is left to your discretion.

Explanation: For a thorough discussion of the blue bottle, consult J. Arthur Campbell's *Why Do Chemical Reactions Occur?* (Prentice-Hall, 1965). Briefly, the mechanism can be explained as follows.

1. The oxygen in the air above the liquid dissolves in the water.
2. The dissolved oxygen reacts with colorless methylene blue to yield blue methylene blue. The

colorless methylene blue acts as a catalyst in this system, because it is consumed and later regenerated.

3. The hydroxide ion in the liquid converts glucose to the glucoside ion plus water. The reaction rate can be increased by increasing the concentration of either the hydroxide ion or glucose (or both).

4. The glucoside ion reacts with the blue methylene blue to yield colorless methylene blue, a hydroxide ion, and other colorless products.

(Step 1) $\text{oxygen}_{(gas)} \rightarrow \text{oxygen}_{(aq)}$

(Step 2) $\text{oxygen}_{(aq)}$ + colorless methylene blue → blue methylene blue

(Step 3) glucose + OH^- ⇌ glucoside ion + H_2O

(Step 4) glucoside ion + blue methylene blue → colorless methylene blue + OH^- + other products

Thus, the system is one in which oxygen in the air reacts with glucose to form colorless products. Even though students are not qualified at this point to understand the mechanism of the reaction, it is a good demonstration to arouse interest. Let individuals pursue the study to the fullest extent you think is fruitful. It can be picked up again later in the course.

Assign for the following day: The study of Experiment 1-1 (pp. 21, 22). Students must prepare their lab record books *prior* to going into the laboratory. Explain that they should write the date, experiment number and title, and prepare a table in which they can record both qualitative and quantitative observations for Part I and Part II when they arrive in class. These should be checked and initialed by the teacher as they go into the laboratory on the following day.

NOTE: Strict attention to prior preparation for experimental work is *essential* early in the course in order to insure good working habits! It is well worth the extra time required to train students in these important details.

Fifth Day. Students will perform Experiment 1-1 individually. If they choose to let the test tube in Part II stand overnight, the results will become more definite.

Assign for the following day: Preparation of formal report on Experiment 1-1, Part I. Stress the importance of part e in the directions for the formal report. Each student is to select *one* hypothesis proposed for an explanation of some observation in Part I and devise plans for testing it in the laboratory. Remind students that *any* hypothesis is good if it meets two criteria: it must be based upon observation, and it must be testable in the laboratory.

Sixth Day. The first part of the period will be devoted to completing observations on Part II of Experiment 1-1. The second part of the period should be devoted to class discussion of the formal report on Part I. Encourage bulletin board displays, posters, creative writing, etc. Check student plans for further investigations on Part I.

Assign for following day: Preparation of lab record book for conducting the further investigations on Part I of Experiment 1-1.

Seventh Day. Laboratory work on individual student plans for further investigations.

Assign for following day: Handing in formal report on further investigations and studying Sections 1-3 through 1-8 in textbook.

Eighth Day. Class discussion on Sections 1-3 through 1-8. During the latter part of the period permit students to go into the laboratory to practice measuring objects in metric units. Meter sticks, graduated cylinders, and balances should be available, along with objects to be measured and weighed.

Assign for following day: The study of Experiment 1-2 and preparation of lab record book for recording observations.

Ninth Day. Perform Experiment 1-2. Note: At this point in the course students may work so slowly that two days could be required to complete this experiment. Adjust schedule to fit your needs.)

Assign for following day: Practice Exercises 1 through 5 at end of chapter (p. 30).

Tenth Day. Discuss results of Experiment 1-2. Go over Practice Exercises in class. Provide remedial work for students having difficulty. Students who do not need remedial work may practice measurements in the laboratory if they feel the need for this.

Assign for following day: The study of Sections 1-9 and 1-10 in textbook.

Eleventh Day. Discuss graphing and clarify questions from students.

Assign for following day: Preparation of lab record book to perform Experiment 1-3, Learning to Work with Variables.

Twelfth Day. Perform Experiment 1-3. Students who wish to do so may perform the optional Part III for extra credit.

Assign for following day: Handing in formal report on Experiment 1-3.

Thirteenth Day. A thorough discussion of the results obtained in Experiment 1-3 is important. Use of an overhead projector is very helpful. The teacher can plot a graph as the discussion proceeds, thus helping to clarify points that may have caused problems.

Assign for following day: Practice Exercise 8 for the students who need more practice in graphing. The study of Section 1-11 "Concept of Density." Encourage creative writing.

Fourteenth Day. Discuss density. Perform the following demonstration.

Demonstration – The Concept of Density

Purpose: To introduce the concept of density with cubes of different materials. Cubes of equal volume and different mass are used in part 1; cubes of equal mass and different volume are used in part 2.

Materials: 1-cm cubes of copper, zinc, iron, lead, cork, sugar, Styrofoam, wax; 500-g cubes of copper, zinc, iron, lead; spheres of Styrofoam and steel having equal volumes; balances; metric rulers.

Metal cubes are used in this demonstration in preference to the usual cylinders or spheres because the majority of students have no difficulty in calculating the volume of a cube ($V = d^3$), whereas if they have to use the formula for the volume of a cylinder ($V = \pi r^2 h$) or the formula for the volume of a sphere ($V = [4/3]\pi r^3$), they can become so involved in the mathematics they lose sight of the fact they are studying density. However, if you choose to work with cylinders or spheres or have to because they are all you have available, you can simplify the determination of volumes by using the following information. The volume of a cylinder that will just fit into a cube or a rectangular solid is 0.79 of the volume of the cube or rectangular solid. A sphere has 0.52 of the volume of the cube in which it will fit.

These metal cubes are not available at most supply houses, but the shop teacher can usually make them. Dimensions will not be so precise as those professionally machined, but they work very well. In fact, the concept of uncertainty in measurement can be stressed if they are not perfectly machined.

Other cubes, such as cork, Styrofoam, wax, sugar, etc., can be used if time permits. Sugar cubes are not exactly 1 cm on a side, but can be used. These cubes serve to let students see how big a 1 cubic centimeter really is.

Part I: Equal Volumes, Different Mass

1. Begin by holding up the cubes of Styrofoam and steel. Ask students which they think is heavier. Since they have had experience with both materials, the answer is pretty obvious to them.

2. Then ask *why* the steel cube is heavier. Do not settle for the answer, "It is more dense." Pursue the question by asking students what they mean by "more dense." Avoid the formal definition of density as "mass per unit volume." Tell them this means nothing to you. Draw them out, and capitalize on such statements as "The particles are closer together" and "There are more particles in the steel cube," or "Each steel particle is heavier than each Styrofoam particle."

3. Measure, or have different students measure, the dimensions of the 1-cm cubes. These should be recorded and the volumes calculated. The volume of a cube = (length)^3.

4. Weigh each cube individually and record. Sample data are given below.

Material	*Volume*	*Mass*	*Density*
copper	1 cc	8.9 g	8.9 g/cc
iron	1 cc	7.8 g	7.8 g/cc
lead	1 cc	11.2 g	11.2 g/cc
zinc	1 cc	7.1 g	7.1 g/cc

Part II: Equal Mass, Different Volumes

1. Cubes of copper, zinc, iron, and lead which all weigh about 500 g are used. Students can see that they differ in volume. They can lift and note the feel of the masses. This is interesting, because the lead seems to be heavier than the iron. Students notice this and it makes a good point for discussion. Why does the lead feel heavier than the iron when we can see it is obviously smaller? Again, it is the density that makes the difference. The mass is packed into a smaller volume.

2. Each cube is measured and its volume calculated. Then the cubes are weighed and the density is calculated by dividing the mass by the volume.

Material	*Length*	*Volume*	*Mass*	*Density*
copper	3.83 cm	56.1 cc	500 g	8.9 g/cc
iron	4.00 cm	64.0 cc	500 g	7.8 g/cc
lead	3.54 cm	44.3 cc	500 g	11.2 g/cc
zinc	4.13 cm	70.4 cc	500 g	7.1 g/cc

Assign for the following day: Preparation of lab record book for Experiment 1-4, Measurement and the Concept of Density. (Note: This experiment requires extensive student preparation *prior* to lab.)

Fifteenth Day. Perform Experiment 1-4.

Assign for following day: Handing in formal report on Experiment 1-4. Encourage creative work.

Sixteenth Day. Discuss results of density experiment. Assign Problems 1 to 4 (p. 27) and Practice Exercises 6 and 7 at end of chapter for students who need remedial work on density.

Assign for following day: The study of Sections 1-12 and 1-13 in textbook.

Seventeenth Day. Discuss the material in Sections 1-12 and 1-13.

Assign for following day: Preparation of lab record book for Experiment 1-5.

Eighteenth Day. Perform Experiment 1-5.

Assign for following day: Handing in formal report on Experiment 1-5.

Nineteenth Day. Discuss the results of Experiment 1-5. Then give the following demonstration quiz.

Demonstration — Effect of Cold Water on Rate of Diffusion of $KMnO_4$

Materials: $KMnO_4$ crystals; ice water; tap water; two 18 mm x 150 mm test tubes; support for test tubes.

Procedure:

1. Tell students you are going to place a crystal of $KMnO_4$ in a test tube of water at room temperature at the same instant one is dropped into a test tube of ice water. Record temperature of each test tube of water.
2. Ask students to write their prediction of results on a sheet of paper and then to record observations as the experiment progresses.
3. Perform the experiment with test tubes in full view of all students.
4. Allow 10 to 15 minutes for the process to develop.
5. Ask students to explain the observed results.
6. Collect the papers at the end of the period.

Results:

1. Much to the surprise of most students, the diffusion process occurs more rapidly in the test tube of *cold* water. This is because heat is being absorbed fairly rapidly from the air, and this causes the molecules near the walls of the test tube to be more active than those in the interior of the test tube, where the water is still cold. As the molecules near the walls acquire energy, convection currents occur. This has somewhat the same effect as stirring, and results in a more rapid distribution of the $KMnO_4$ into the cold water.
2. During discussion of results, ask how one could test whether the diffusion in cold water would be affected if the water could not absorb heat near the walls of the test tube. The answer is to place the ice-water test tube in an ice-water bath to keep its temperature constant. This makes a good investigation for creative work.

Assign for following day: Preparation of lab record book to do Exercise 1-1, Learning to Use a Chemistry Handbook.

Twentieth day: Let students do Exercise 1-1 in class. If there are not enough copies of handbooks available for each student, permit them to work in small groups.

Assign for following day: Handing in Exercise 1-1; review of the chapter as needed. Encourage students to do creative work. Suggestions are given at the end of the chapter, but students may have better ideas.

Twenty-first day: Discuss the use of the handbook; assist students who have had difficulty in locating information; clarify any concepts on which students have questions. Allow time for creative work, both laboratory work and other types.

Assign for following day: Answering questions in self-test at end of Chapter 1.

Twenty-second day: Help students who are having difficulty with items of the self-test. Continue creative work.

Assign for following day: Test on Chapter 1. Remind students to study the objectives stated at the beginning of the chapter and to use the practice exercises and self-test as a very effective way to prepare for the "test which counts."

Twenty-third day: Give test on Chapter 1. Students who score at least the minimum number of points on the Chapter 1 test should look over Chapter 2 and prepare a tentative daily schedule of activities. It is extremely important that they plan carefully so that they can maintain a pace profitable for them. Insist that each student make a schedule; point out to them that, later, it can be revised as needed, but the initial, formalized organization of their time is essential to their good progress.

Those who need remedial work prior to taking a second test should be given proper assistance.

EXPERIMENT 1-1
PRACTICE IN SCIENTIFIC OBSERVATION

Purposes:

To provide students with an opportunity to make observations, both qualitative and quantitative, so they become aware of the value of quantitative data; to provide opportunity for students to distinguish between observation and inference, or interpretation; to introduce the technique of proposing a hypothesis as a possible explanation for what has been observed; and to provide opportunity for students to plan an investigation of their own to test a hypothesis they have proposed.

Materials: (per student)

Part I:

1 250-cc beaker
2 sugar cubes

Part II:

1 100 mm x 13 mm test tube
1 20-cm length of bare copper wire
5 cc 0.2M $AgNO_3$ (34 g/liter)

Possible Results:

Part I:

A sugar cube usually dissolves completely in tap water at room temperature in about 20 minutes. During this time, bubbles are seen to come from the cube and rise to the top of the water. Gradually, some of the crystals crumble off and fall to the bottom of the beaker. At the end of 15 to 20 minutes, when the student picks up the beaker and tilts it gently to and fro, the more dense layer of solution in the lower region of the beaker is quite evident.

Quantitative observations will include: the *rate* at which bubbles rise (one every two seconds, etc.); the time for dissolving to be complete; the time elapsed prior to obvious crumbling of the cube; etc. Qualitative observations will include: bubbles escaped from the sugar cube; the cube crumbled; there is a different layer of liquid on the bottom; etc.

Inferences, or interpretations, will include such statements as: the bubbles are air; the rate of dissolving slowed near the sugar cube because that portion of the water held more dissolved sugar; the thick layer in the bottom was syrup; etc.

One hypothesis might be: air was trapped in the cube and was driven out as water penetrated the cube. To test this, a student might crush a sugar cube in a mortar and pestle and repeat the experiment.

Part II:

Within a few minutes silver crystals begin to form on the copper wire. Students usually describe this as "moss." As the reaction proceeds, the colorless solution becomes faintly blue. After the wire has stood in the solution overnight, the blue color is quite noticeable.

If possible, have a binocular microscope available so students can remove the copper wire and observe the silver crystals. A good magnifying glass is a barely acceptable substitute for the microscope.

At this point most students will not be able to explain the changes taking place. Do not tell too much. Encourage them to propose hypotheses and to test these in the laboratory. Some of these will make excellent free labs. Encourage reading in other textbooks as students search for explanations. They will return to this activity several times during the course.

Help students to distinguish between an observation and an inference. At this stage, some students tend to confuse the two.

Postlab discussion should enhance the importance of careful observation, the value of quantitative data, the difference between observation and inference, the possible techniques for proposing a hypothesis, and the importance of devising further experiments to test a hypothesis.

This is a good time to stress the importance of creative thinking in science and to let the students know that extra credit will be given for free labs if a formal report is submitted.

EXPERIMENT 1-2
SOME USEFUL LABORATORY TECHNIQUES

Purposes:

To provide an opportunity for students to learn some useful and timesaving techniques; to introduce the test for oxygen and hydrogen gases; and to teach the meaning of a flow chart and how to use it.

Materials: (per student)

1 10-cc graduated cylinder
1 50-cc graduated cylinder
1 pipet
1 50-cc beaker
4 100 x 13 mm test tubes
5 cc 3% hydrogen peroxide (Note: Do NOT use more concentrated solutions as they are quite dangerous.)
5 cc 3M HCl
½ teaspoon baking soda, $NaHCO_3$
a "pinch" of MnO_2
5 cc limewater (saturated solution of $Ca(OH)_2$ that has been filtered to remove the excess solid)
5 to 10 wood splints
1 small piece of mossy zinc
0.5 g $Pb(NO_3)_2$
10 cc 0.1M KI solution (1.6 g/liter)

Prelab Discussion:

Demonstrate the technique for collecting and testing for oxygen and hydrogen gases that the students will use in this experiment. Stress the importance of keeping the hydrogen test tube mouth downward.

Safety precautions must be emphasized because of the corrosive properties of acid, the explosive nature of hydrogen and oxygen mixtures, and the necessity for wearing eye protectors in the laboratory.

Go over the meaning of the symbols in the flow chart. Encourage students to make a flow chart for each future experiment, where it is appropriate.

Postlab Discussion:

No formal report is needed for this experiment because it is a learning activity and not a discovery one. Student questions may be discussed for clarification. Emphasize the usefulness of being able to estimate volumes of liquids when precise measurement is not needed. This saves time and dishwashing.

EXPERIMENT 1-3
LEARNING TO WORK WITH VARIABLES

Purposes:

To introduce the technique of graphing linear relationships from data obtained in the laboratory; to provide opportunity to learn how to determine the slope of a line; and to provide practice in interpreting graphical information.

Materials: (per student)

1 10-cc graduated cylinder
balance to weigh ± 0.01 g
5 can lids, or other circular objects of varying size
string and metric rule, or metric tape measure

Results:

Part I:
An increasing linear relationship is obtained between the volume of water and its corresponding mass. The slope of the line should be 1.0 gram per cc of water. This is the density of water. Some students may recognize this if they have had physical science. If they do not, it may be pointed out later when density is studied.

Analysis of Data, Part I:

1. Mass will go from about one gram up to about 10 grams.

2, 3. The independent variable is volume in this case; the dependent variable is mass. The graph should be a straight line, running through the origin; the data points should fall close to, if not on, the line. If any points deviate severely from the line, the student should check his data, and possibly do another measurement.

4. The slope is the change in mass divided by the change in volume, for any two points on the line. It should be close to 1.0 g/cc.

5. Since this graph goes through the origin,

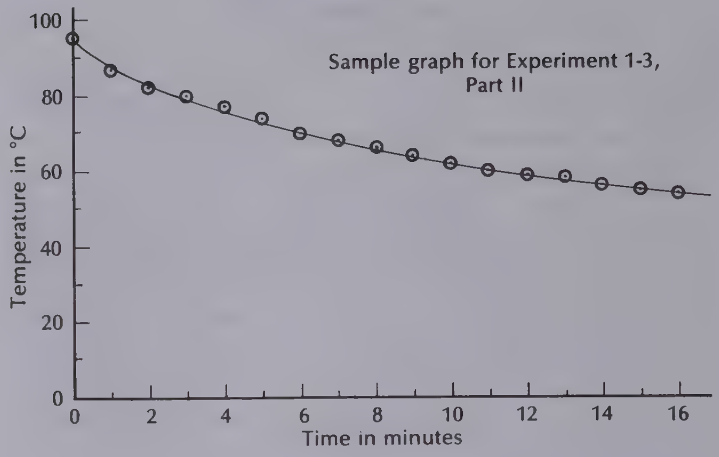
$$\text{slope} = \frac{\text{mass}}{\text{volume}}$$

6. At zero volume the mass is zero.

7. The average increase in mass per cc of water is 1 gram.

8. Since mass = slope x volume, the predicted mass of 100 cc of water is 100 grams.

Analysis of Data, Part II:



1, 2. See sample graph. The independent variable is time, the dependent variable, temperature. The graph may drop more steeply at the beginning than at later times.

3. First minute, temperature change was 7.5°C, so ratio is 7.5°C/min. Eighth minute, change was 2.0°C, so ratio is 2.0°C/min. Fifteenth minute, change was 0.5°C, so ratio is 0.5°C/min.

4. An astute student may suggest that as the temperature of the water approaches room temperature, the rate of cooling will gradually slow and approach zero.

Part III:
Although this experiment has no relationship to chemistry, it is important because it provides further practice in graphing data and interpreting the graph. Also, it provides direct information for the meaning of pi, π. Most students have had no experience with its derivation and are surprised to find that the value they have accepted and used for so long is really the ratio of the circumference of a circle to its diameter.

Analysis of Data, Part III:

1, 2. Diameter, the independent variable, is plotted on the horizontal axis. Circumference, the dependent variable, is plotted on the vertical axis. The graph is a straight line which, if extended, would pass through the origin.

3. The ratio of circumference to diameter is π (approx. 3.14). Note: Students usually obtain values close to this but their values will depend upon the care with which they measured the variables.

4. Circumference/diameter = 3.14.

5. Yes, because a circle with zero diameter would have zero circumference.

6. The *predicted* value is asked for. Therefore, the answers will depend upon the students' values for the slopes of their own graphs. If they have done careful work, the answer should be close to 78 cm.

Graphs should be checked carefully to be sure students are using the proper axes, labelling the parts correctly, and are indicating the scale of units used. Attention needs to be given to drawing a "smooth" line. The line need not always be straight, but neither should it be "wavy." Stress the fact that because of uncertainties in measurements, it is highly unlikely that all points will lie exactly on any line, and that this is to be expected.

EXPERIMENT 1-4
MEASUREMENT AND THE CONCEPT OF DENSITY

Purposes:

To provide practice in using instruments for making measurements; to develop skill in the use of significant figures as a method of expressing uncertainty in measurement; and to provide experience in determining the density of objects by different methods.

Materials: (per student)

1 regular object (this may be a cube or cylinder of metal from a specific gravity set)
1 irregular object (a small piece of roll sulfur, small samples of minerals of known density, irregular pieces of pure metals of known density, or stones picked from a stream bed)
1 object which floats in water (a small piece of cork or wood of known density may be used)
1 graduated cylinder of proper size
1 balance
1 metric rule
Handbook of Chemistry and Physics

Results:

Density is a difficult concept and extensive experience with objects of varying density is needed to develop an understanding of the concept. It is hoped that students will do sufficient reading prior to the experiment to prepare them to determine the volume of the objects by (1) direct measurement of the sides of regular objects and calculation of the volume by use of formulas and (2) displacement of water by fully submerged objects.

During the prelab discussion, students should be encouraged to give thought to how they might determine the density of the various objects, and to enter in their record book their predictions about the densities of the different objects prior to the investigation.

Hopefully, the percentage error will not exceed 5%; however, if students have had little experience in measuring, the error may be greater. It is unwise to grade on percentage of error, because this leads to manipulation of data. Instead, increased accuracy may be encouraged by suggesting to students with a great percentage of error that they should repeat the experiment in an effort to improve the quality of their work.

Analysis of Data:

1. Stress the importance of using the proper number of significant figures.
2. We recommend that each student use the handbook to look up the density of his objects, because this enables him to become more familiar with this important reference tool. However, if it is more convenient for you, these values may be placed on the bulletin board.
3. Students will be asked to calculate these kinds of errors in many experiments, so it is important that they learn the proper procedure. (See pp. 19 and 20 of student text.)
4. Answers will vary. The hypothesis is the point to be stressed here.
5. Answers vary. If the measurement is done precisely (as with calipers) and the object is quite regular, the use of linear measurements will be the most precise. With irregular objects volume is best measured by the water displacement method. But in any case, the measuring must be done with great care.
6. They should all have approximately the same value for density, because density is mass per unit volume and would be the same for any given material regardless of the volume the student is using. The concept that the mass-to-volume ratio is constant for all samples of the same substance is extremely important. Check the students' answers carefully because they provide evidence about how well each student understands the concept of density.
7. The density of water equals the slope of the graph. It should be 1.0 g/cc.

EXPERIMENT 1-5
HEAT ENERGY AND RANDOMNESS

Purposes:

To introduce the concept of randomness as influenced by heat energy; to provide further practice in making predictions; and to stimulate creative thinking by use of the follow-up demonstration quiz (p. 14 of this guide).

Materials: (per student)

2 tiny crystals of potassium permanganate, $KMnO_4$
1 100-cc beaker
2 18 x 150 mm test tubes
burner
ringstand, ring, wire gauze
beaker tongs
timer
thermometer

Prelab Discussion:

Demonstrate the proper method of pouring hot water. Caution students about careless handling of boiling water. Remind students to use *very tiny* crystals of potassium permanganate and that the crystals need to be the same size.

Postlab Discussion:

Compare student results and discuss reasons for their findings. Discuss the meaning of randomness and what is meant when we say that the particles of a solid are more random when the solid is dissolved in a liquid than when it is a crystal. There are more different ways for the particles in the solution to be arranged (with no outward change in appearance of the solution).

The importance of heat energy in this experiment should be emphasized. Let students discuss their predictions in Question 5, but give no clues about their correctness, as this sets the stage for the demonstration quiz to follow.

Questions:

1. In the one containing hot water.
2. In the one containing hot water.
3. It appeared to have increased the rate of dissolving.
4. Heat appeared to have favored randomness.
5. Prediction: it would dissolve more slowly. This prediction follows from the results of this experiment. However, as the demonstration quiz shows, the crystal will actually dissolve more rapidly unless special precautions are taken.

EXERCISE 1-1
LEARNING TO USE A CHEMISTRY HANDBOOK

Purposes:

To introduce the student to the *Handbook of Chemistry and Physics* by asking him to look up information in those sections which he will find most useful.

Materials:

In addition to the *Handbook of Chemistry and Physics,* several other handbooks should be available, preferably from different publishers. This will provide an opportunity for the student to become familiar with more than one handbook. It is very important that the title, authors or editors, publisher, and publication date or edition be entered in the laboratory record book, because the handbooks vary. Some examples of handbooks are:

Aylward, G. H., and T. J. V. Findlay, ed., *Chemical Data Book,* John Wiley & Sons, second edition, 1966 (paper).

Aylward, G. H., and T. J. V. Findlay, ed., *S. I. Chemical Data,* John Wiley & Sons, 1971 (paper).

Handbook of Chemistry and Physics, The Chemical Rubber Co., Cleveland, revised yearly.

Kaye, G. W. C., and T. H. Laby, *Table of Physical and Chemical Constants,* John Wiley & Sons, 13th edition, 1966.

Summers, D. B., *The Chemistry Handbook,* Willard Grant Press, Boston, 1970.

Prelab Discussion:

Some discussion of the importance of a handbook is in order. Encourage students to become familiar with the index of the handbook and how to use it so that they may make the most of this source of tabulated information.

Answers to Questions:

(Taken from the *Handbook of Chemistry and Physics,* The Chemical Rubber Company, Cleveland, 1966)

	Answer	Page No.
a.	0.51 A°	a. F-124
b.	Remove to fresh air and call for pulmotor; apply artificial respiration for at least 1 hour or until pulmotor arrives; the inhalation of ammonia is often of value.	b. xxv
c.	Cronstedt, 1751	c. B-123
d.	6.38 grams	d. B-210
e.	$Pb_2Fe(CN)_6 . 3H_2O$	e. B-187
f.	0.923 at 15°C	f. C-98
g.	290°C (decomposes)	
h.	Vapors toxic; violent explosion with sodium fires	h. xxvii
i.	8.90	i. A-143
j.	Dissolve 0.5 g congo red in 90 ml distilled water and 10 ml alcohol.	j. D-63
k.	0.0005 angstrom	K. E-126
l.	The solid which separates when a saturated solution freezes. It contains the solvent and the solute in the same proportion as they were in the saturated solution.	l. F-64
m.	1426×10^6 km	m. F-117
n.	Multiply by 10	n. F-170

CHAPTER 2

CLASSIFICATION OF MATTER

Goals:

1. To introduce students to the meaning of physical and chemical properties of matter through laboratory experiences.
2. To provide an opportunity for students to observe changes in state and to consider energy changes that are associated with changes in state.
3. To study a typical metal (iron) and a typical nonmetal (sulfur) as elements, then to observe a chemical change as they combine to form the compound iron sulfide.
4. To provide a background for understanding attractive and repulsive forces in a way that can be applied to the study of chemical bonding.

Teaching Hints:

1. In the discussion of physical and chemical changes, reference is made to the fact that some modern chemists believe that if the physical properties of matter change, then a chemical change has occurred. For further discussion of this modern concept refer students to the *Encyclopaedia Britannica*, 1972 edition, under the topic "Chemical Reactions."
2. Experiment 2-1 may be frustrating to students. This is deliberate planning on the part of the authors in order to point up the fact that there is no easy way to distinguish elements, mixtures, and compounds. To do so with confidence requires extensive, and sometimes complicated, testing in the laboratory. As the students become aware of this, you can help them to see beyond their frustrations to the real meaning of the experiment.
3. Continue to initial lab record books before permitting students to perform an experiment. It is important that they study the experiment so thoroughly *prior* to going into the laboratory that they know what they are going to do and what they should accomplish. Only if they study the experiment in detail — from the introductory remarks through the questions to be answered at the completion of the experiment — will their laboratory work be meaningful. Otherwise, it is simple "cookbook" chemistry devoid of real meaning to the students.

   This requires *constant* diligence on the part of the teacher. If students do not know how to prepare appropriate data tables, take time to assist them early in the course. Later experiments require more detailed and complicated data tables, so the time to learn this technique is now.
4. Check carefully to be sure that everyone has a time schedule and that he is following it within reasonable expectations.
5. Formal reports on experiments should be checked *very* carefully *early* in the course so that students form proper reporting habits.
6. Encourage creative work of *all* types. Put a sheet of paper on the bulletin board so that students can sign up to do bulletin board displays. Provide a space for name of student, the date they wish to have the bulletin board, and the topic it will cover. We suggest that each student be allowed to keep his display up for one week.
7. It is not necessary for students to memorize the symbols for all elements. However, you should have them check the names of the ones you will require.
8. Draw the students' attention to the photographs of food labels on pp. 58 and 59. See if they can guess what common foods bore these labels.

EXPERIMENT 2-1
CLASSIFICATION OF MATTER

Purposes:

To introduce different methods for classifying matter through direct experience; to acquaint students with the difficulty of classifying matter by inspection; to provide a background upon which to build the more subtle concepts involved in adequate classification of matter; to provide experience in proposing hypotheses and in devising suitable tests for them.

Materials: (per student)

15 covered vials 4 cm high by 2 cm in diameter, containing the following samples of matter, one sample per vial. Vials are coded as follows:

A. roll sulfur
B. chips of granite
C. table sugar
D. mixture of iron filings and sulfur
E. copper sulfate solution
F. white sand
G. strip of aluminum foil
H. mixture of sawdust and sugar
I. vinegar
J. sodium hydroxide solution (0.1M)
K. charcoal chips
L. distilled water
M. potassium permanganate solution
N. NaCl, solid
O. a few strips of neutral litmus paper

This list corresponds with the illustration in the laboratory manual. Add to the list if you wish. Iron filings, lead foil, a vial of air, powdered milk, and any number of other things can be used.

Prelab Discussion:

During the prelab discussion, caution students about spillage, since some of the samples may be corrosive to hands and books. Instruct them to use the *letter* on each vial for classification purposes. Each vial should be labeled with the above letters, but the name of the sample should not be given to students.

Results and Postlab Discussion:

Part I

1. *Homogeneous matter*	*Heterogeneous matter*
A. sulfur	B. granite
C. sugar	D. Fe + S
E. $CuSO_4$ solution	H. sawdust + sugar
F. white sand	
G. Al foil	
I. vinegar	
J. NaOH solution	
K. charcoal	
L. H_2O	
M. $KMnO_4$ solution	
N. NaCl	
O. litmus paper	

2. *Solid*	*Solution*	*Unable to determine*
A. sulfur	E. $CuSO_4$	I. vinegar
C. sugar	J. NaOH	L. water
F. sand	M. $KMnO_4$	
G. Al		
K. charcoal		
N. NaCl		
O. litmus paper		

3. There will be some variety of answers. Actually, one cannot determine whether matter is an element or a compound by inspection. So long as students have a *reason* for their classification, answers should be accepted. One purpose of this activity is to provide evidence that classification of matter is not a simple process.

3. *Element*	*Compound*	*Unable to determine*
A. sulfur		C. sugar
G. Al		F. sand
K. charcoal		N. NaCl
		O. litmus paper

4. *Metal*	*Nonmetal*
H. Al	A. sulfur
	C. sugar
	F. sand
	K. charcoal
	N. NaCl
	O. litmus paper

5. *Acid*	*Base*	*Unable to determine*
E. $CuSO_4$	J. NaOH	L. water
I. vinegar		M. $KMnO_4$

Part II

1. Answers will vary. Experimentation should be encouraged.

2. Answers will vary, but again experimentation should be encouraged.

Questions:

1. Answers vary. Color, density, solubility in water, and other criteria could be used, depending upon the purpose of the experimenter.

2. (a) Matter may appear to be composed of only one element when in reality it may be composed of two or more elements chemically combined. Sugar appears to be an element, but heating will show this is not true.

(b) Homogeneous liquids might contain more than one substance. The only way to find out is to conduct the proper tests.

3. Answers will vary. Examples: a. air b. salt and pepper c. ice d. alcohol e. oxygen

The point to emphasize during the postlab reporting is that precise classification of matter is not easy. One must know how to test and what to look for if classification systems are to be refined. The variety of student answers will point up this fact. The stage is set for the textbook discussion of what the chemist considers to be an element, a compound, or a solution.

NOTE: In grading this experiment, give large amounts of credit to those students who realize that matter cannot be classified by inspection only. Do not lower the grade if the student fails to classify the samples correctly. The most credit should be given to those who classify the majority of the samples as "Unable to determine." On the basis of the evidence available this is reasonable.

EXPERIMENT 2-2
PHYSICAL AND CHEMICAL PROPERTIES OF MATTER

Purposes:

To introduce the student to operational definitions of physical and chemical properties; to obtain some evidence of chemical change; to provide opportunity for the student to propose hypotheses to explain his observations.

Materials: (per student)

6 cm x 4 cm piece of aluminum foil
10 cc 6M HCl (513.0 cc concentrated acid per liter of solution)

10 cc 6M NaOH (240 g solid NaOH per liter of solution)
3 cc 1M ammonia water (66.3 cc concentrated reagent per liter of solution)
8 cc distilled water
scissors
steel wool
metric ruler
balance, readable to 0.01 g
evaporating dish

Prelab Discussion:

Caution students about handling HCl and NaOH. Remind them to review the technique for collecting and identifying common gases (see Experiment 1-2).

Results and Postlab Discussion:

1. The rubbed foil usually reacts faster than the unrubbed. This is because a thin coating of aluminum oxide protects the Al from the reagent and is partially removed by rubbing. Some students do not rub vigorously enough; others inadvertently re-coat the rubbed foil by careless handling.

2. Call attention to the fact that the reactions can be represented by word equations:

aluminum + hydrochloric acid →
hydrogen + aluminum chloride

The aluminum chloride is the residue left in the evaporating dish after all liquid has disappeared. (NOTE: Frequently this residue is yellow—probably due to impurities.) The other word equation is

aluminum + sodium hydroxide →
hydrogen + sodium aluminate

3. Discuss the formation of a precipitate upon addition of ammonia water (NH_4OH) to the aluminum chloride solution. The gelatinous solid is characteristic of aluminum precipitates.

aluminum chloride + ammonium hydroxide →
aluminum hydroxide + ammonium chloride

(NOTE: If word equations are used routinely and without fanfare to represent reactions, students become accustomed to them, and this makes the transition into writing chemical equations quite natural.)

Questions:

1. Testing for color, mass, and odor causes no change in the composition of matter, whereas testing for the ability to react with acids and bases causes the composition of the matter to change.

2. Evolution of a gas; disappearance of the aluminum foil; evolution of heat (the test tubes became warm); and formation of a precipitate.

3. Suggest that students do some research reading on the properties of aluminum prior to answering this question. The reactions may differ because usually (not always) there is a protective coating of aluminum oxide on the surface of aluminum metal.

4. Hydrogen, because it "popped" or "barked" when ignited. This corresponds to the results obtained in Experiment 1-2.

Questions to Ponder:

Students are not expected to give the "correct" answer to these questions at this time. Give credit if they appear to have done serious thinking. Do not let these answers lower their grade on this experiment unless they have omitted them. In that case, give an "incomplete" and ask students to make their best guess. We consider this activity an important part of the chemistry course.

1. Probably, but this would require experimentation.
2. If students think about the composition of the two reagents, they should be able to guess that the hydrogen must come from the acid.
3. Energy must be released during the reaction.
4. Aluminum chloride.
5. Aluminum hydroxide.

EXPERIMENT 2-3
MELTING AND FREEZING OF SOLIDS

Purposes:

To provide quantitative information on the change in state of a pure substance; to provide experience in recording and graphing data; to provide experience in interpreting data; to relate the observations and data to the opposing tendencies of minimum energy and maximum randomness.

Materials: (per pair of students)

15 g of paradichlorobenzene in a 18 x 150 mm test tube
2 thermometers (range −10°C to 110°C)
60 cc gas collecting bottle
250 cc beaker
timer or watch with sweep second hand
burner
ringstand, ring, asbestos mat
buret clamp

Prelab Discussion:

About 10 to 15 g of paradichlorobenzene crystals can be weighed and placed in a test tube (18 x 150 mm). The other tubes can be filled visually, using the first as a guide.

The temperature of the water is not critical for the cooling curve. However, for the melting curve, the water temperature must be between 85°C and 90°C when the test tube of paradichlorobenzene is inserted. If it is cooler, there is too little heat for complete melting. If it is warmer, the plateau region in the curve is lost.

Caution students about not overheating the solid when melting. Some instruction may be needed on the proper technique for reading a thermometer. The lines of sight from eye to thermometer should be perpendicular to the mercury column for accurate reading. It is important that proper data tables be prepared *prior* to the laboratory period. If students are inexperienced in laboratory work, they may need assistance in setting up the data tables. (See sample in this discussion.)

Results and Postlab Discussion:

The care with which students work will determine how well the plateau region shows on the paradichlorobenzene graph. This is a crucial part of the experiment. A good graph should be shown for discussion purposes. The relatively constant temperature during change of state of pure substances is a key concept. Let students propose hypotheses about why this is true. It is recommended that the answer not be given at this time. Encourage investigation and thinking.

If students repeat the experiment with an unknown, they may use candle wax or paraffin. The graph for either wax will not show a plateau because these mixtures contain substances of varying melting points. It may be pointed out that one of the major differences between a compound and a mixture is that in mixtures the individual components retain their own properties. Some *pure* substances which could be used as "unknowns" include:

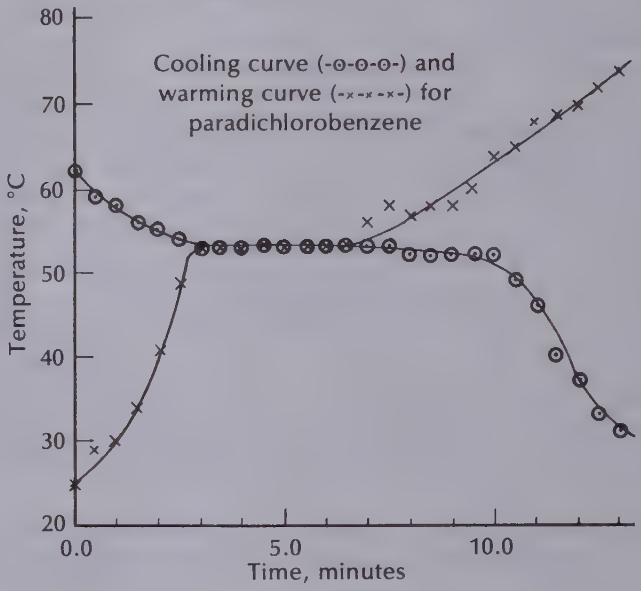
lauric acid (melting point = 44° C)
octadecanol (melting point = 58.5° C)
palmitic acid (melting point = 65° C)
acetic acid (melting point = 16.6°C) Note: This substance tends to supercool.

SAMPLE DATA TABLE FOR FOR PARADICHLOROBENZENE

Time, minutes	*Temperature, °C (cooling)*	*Temperature, °C (warming)*
0.0	62	25
0.5	59	29
1.0	58	30
1.5	56	34
2.0	55	41
2.5	54	49
3.0	53	53
3.5	53	53
4.0	53	53
4.5	53	53
5.0	53	53
5.5	53	53
6.0	53	53
6.5	53	53
7.0	53	56
7.5	53	58
8.0	52	57
8.5	52	58
9.0	52	58
9.5	52	60
10.0	52	64
10.5	49	65
11.0	46	68
11.5	40	69
12.0	37	70
12.5	33	72
13.0	31	74

Analysis of Data:

1. See sample graph.



2. Any temperature between 51°C and 55°C is considered a good value. Students will note some variation of temperature within the plateau on any individual graph, but the average value will be within these figures for the majority of students. It is hard for some students to accept the idea that the melting point and the freezing point for a substance are identical. Experimentation with ice and water is to be encouraged.

3. 53°C

4. The plateau signifies that there was no change in the temperature of the paradichlorobenzene during the processes of melting and freezing. During

melting, the heat energy being supplied is used to unpack the particles of the solid. When these particles re-assemble during solidification (freezing), energy is released.

5. Solids exhibit order — they maintain a constant shape; the particles are not free to move around. As solids melt, heat energy is absorbed and the particles move apart, becoming more randomly arranged. Thus, both the total energy and the amount of randomness are greater in the liquid than in the solid.

Demo

EXPERIMENT 2-4
ELEMENTS, MIXTURES, AND COMPOUNDS

Purposes:

To familiarize students with the properties of two elements, iron and sulfur; to provide experimental evidence that elements differ from each other in several properties; to provide experience in making a compound from elements; and to compare the properties of a compound with those of the elements from which it is formed.

Materials: (per student)

10 g iron filings
6 g crushed roll sulfur
15 cc 3M HCl
6 13 x 100 cc test tubes
1 18 x 150 cc test tube
1 horseshoe magnet
balance, precision of 0.01 g
1 250 cc beaker

Prelab Discussion:

1. Remind students to check the "Laboratory Techniques and Procedures" section of the laboratory manual to be sure they understand how to check for a dissolved solid.

2. Caution students about handling HCl.

3. Stress the fact that this experiment is designed to enable students to learn through experience how elements may be distinguished from their compounds.

4. In Part III, step 4, it is important to heat the test tube of elements in the burner flame until an evident glow is seen inside the tube. Remind students to remove the test tube from the flame, wait until the glow is no longer visible, then plunge the test tube into the beaker of water.

5. If the sulfur catches fire in Part III, step 4, do not become alarmed. Place an asbestos mat over the mouth of the tube to extinguish the flames, or simply remove the test tube from the flame for a few seconds.

Questions:

1. *Mixtures*	*Compounds*
a. The components of a mixture do not lose their individual properties.	a. The components of a compound lose their individual properties.
b. The parts of a mixture can be present in a variety of proportions.	b. The parts of a compound are always present in the same definite proportion by weight.
c. Mixtures can be separated by physical procedures.	c. Compounds can be separated only by chemical procedures.
d. Mixtures form without release of energy in the form of light or heat.	d. When compounds form, there is usually evidence of released energy in the form of light, heat, or both.

2. A gas was produced when iron reacted with hydrochloric acid and also when iron sulfide reacted with hydrochloric acid. The gas produced by the reaction of iron and hydrochloric acid was hydrogen. It "popped" or "barked" when ignited in the presence of air. The gas produced when iron sulfide reacted with hydrochloric acid was not hydrogen. It had a foul odor and did not "pop" or "bark" when a glowing splint was brought to the mouth of the tube.

(NOTE: The latter gas is hydrogen sulfide, H_2S. Students may not have the background to identify it, but this could make an excellent free lab if a student is carefully supervised. Hydrogen sulfide is extremely poisonous — more so than hydrogen cyanide — so work should be done in a fume hood, near an open window, or out-of-doors.)

3. The ideal answer is no; however, actual answers will vary. Grade mostly on the thought that can be offered to justify an answer. Encourage a free lab to answer the question.

4. This is a difficult question. The best guess for students at this point is that energy is being given off in such large quantities that both light and heat are evolved. They may not know now that the union of these two elements results in a compound that has much less potential energy than the separate elements. Refer back to this as you study chemical bonding.

5. Just as the heat of friction is necessary to start a match burning, so some heat energy is needed to cause the elements to begin to react. This is formally known as activation energy. Do not stress this unless you have students who are capable and interested in going into the full story. By all means, encourage as much investigation as you can!

6. Do not assign this question on a "you-have-to-do-it" basis. Instead, invite students to try it if they wish.

a. Percentage of iron in iron sulfide:

7 g ÷ 11 g = 0.64 or 64%

b. Percentage of sulfur in iron sulfide:

4 g ÷ 11 g = 0.36 or 36%

c.

Element	Percentage in iron sulfide	÷	atomic weight	=	relative number of atoms (round off)
iron	64	÷	56	=	1
sulfur	36	÷	32	=	1

Calculated simplest formula for iron sulfide = FeS

Encourage all students to try the question. They will be asked to perform an experiment later (Exp. 9-6) in which they will determine the formula for an unknown compound in the laboratory.

EXPERIMENT 2-5
FORCES BETWEEN LIKE AND UNLIKE MAGNETIC POLES

Purposes:

To introduce students to the concept of force acting at a distance by means of magnets which they can see and "feel," and to provide further opportunity for the graphical analysis of quantitative data.

Materials: (per pair of students)

1 ringstand
1 meter stick or metric ruler
masking tape
platform balance (platform must not respond to magnet)
wood splints
1 pair bar magnets (*Note:* When not in use, the two magnets should be stored with opposite poles side by side.)
lump of modeling clay
1 buret clamp

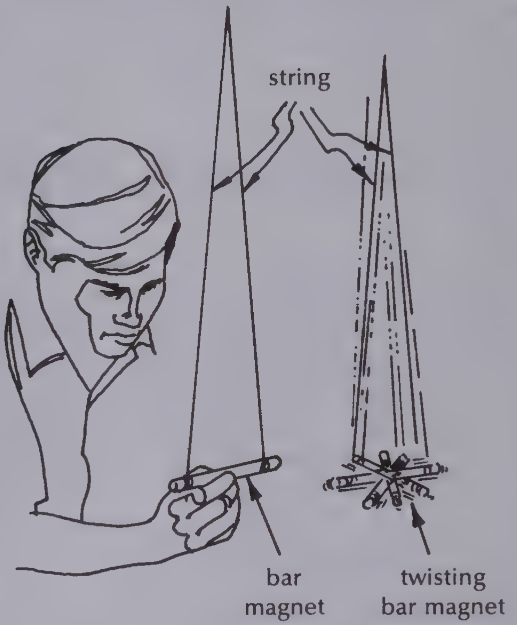
Prelab Discussion:

Stress the necessity of having a prepared data table in the laboratory record book prior to class. The data table should include three columns for LIKE poles and three columns for UNLIKE poles. For each position of the top magnet, both the distance between magnets in cm and the apparent weight of the bottom magnet and the clay in g will be recorded. The third column is for the calculation of the magnetic force in g.

It is important that the actual weight of the bottom magnet and the clay be recorded. Calculate the force in grams for each reading by determining the *difference* between the actual weight of the magnet and clay and the apparent weight. These are the force values which will be used in plotting the graph.

Results and Postlab Discussion:

It is important to discuss the *meaning* of the results of this experiment. Since the magnetic forces of attraction and repulsion can be "felt," when the poles are at a distance, the idea of a force field can be presented in a very real sense. Changes in position of an object affected by a field may be accompanied by changes in both potential and kinetic energy.



Discussion can be centered around a demonstration in which two bar magnets are suspended in stirrups attached to an overhead support (see illustration). Moving one magnet will cause the other to respond. When you discuss magnetic fields, it might be well to make students aware that the earth's gravitational field is also affecting the magnets.

These general concepts of force fields and energy are fundamental to understanding chemical bonds, so spend time making sure that students comprehend the concepts illustrated by magnetic fields.

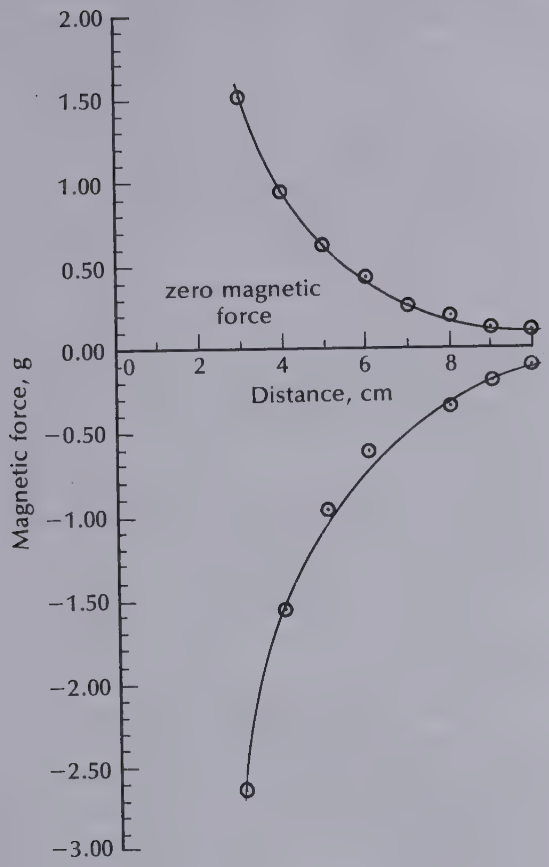
Analysis of Data:

SAMPLE DATA TABLE

LIKE POLES FACING			UNLIKE POLES FACING		
Distance Between Magnets, cm	Apparent Weight, g	Magnetic Force, g	Distance Between Magnets cm	Apparent Weight, g	Magnetic Force, g
10	56.89	0.10	10	56.68	−0.11
9	56.90	0.11	9	56.58	−0.21
8	56.97	0.18	8	56.43	−0.36
7	57.04	0.25	7	56.37	−0.42
6	57.21	0.42	6	56.15	−0.64
5	57.40	0.61	5	55.81	−0.98
4	57.72	0.93	4	55.21	−1.58
3	58.30	1.51	3	54.12	−2.67

Actual weight of magnet and clay = 56.79 g

Note: Magnetic force = apparent weight minus actual weight.



Questions:

1. Like poles cause the apparent weight to be greater than the actual weight and thus give positive values for the magnetic force. Unlike poles cause the apparent weight to be smaller than the actual weight and thus give negative values for the magnetic force.

2. When the force values were negative, the force should be labeled as attractive because the lower magnet was being pulled *toward* the upper magnet.

3. The force increases rapidly as the distance between the magnets decreases. The theoretical mathematical expression is the inverse square law: the force is inversely proportional to the square of the distance.

4. Potential energy is position energy. If the poles are moved against the direction of the force, work is done and the potential energy is increased. If the poles are allowed to move in the direction of the force, the potential energy decreases and kinetic energy (energy of motion) is apparent.

CHAPTER 3

PROPERTIES OF LIQUIDS AND THE SOLUTION PROCESS

Goals:

1. To extend understanding of the concept of dynamic equilibrium in different systems.
2. To emphasize the unique properties of water.
3. To introduce the student to the different kinds of solutions and the characteristics of solutions in general.
4. To provide an understanding of the mechanism of the solution process.
5. To introduce the mole concept as a convenient method for preparing solutions of known concentration.
6. To provide laboratory experience with energy changes that accompany the solution of solids.
7. To provide additional experience in interpreting data.

Teaching Hints:

1. One of the continuing themes of KEYS TO CHEMISTRY is the concept of equilibrium. Experiment 3-1 seeks to provide laboratory experiences that will make this concept more meaningful to the student. There are several demonstrations that can be done to emphasize the effect of the vapor pressure of a liquid upon its rate of evaporation.

A. Place 10 drops of each of the following liquids on separate wads of cotton: water, alcohol, and carbon tetrachloride. Ask students to predict which will evaporate most rapidly when the wads of cotton are swished across the chalkboard. Then, one at a time, swish the wet cotton wads across the chalkboard and record the time required for evaporation. Compare the time for evaporation with the vapor pressure of the liquids at room temperature.

Results:	*Vapor Pressure (20°C)*
Most rapid: CCl_4	76.7 torr
Second: alcohol	43.9 torr
Last: water	17.5 torr

B. Place 5 drops of each of the following liquids on separate watch glasses: water, alcohol, carbon tetrachloride, and acetone. Record the time required for each to evaporate completely.

Results:

Liquid	*Time*	*Vapor Pressure (20°C)*
acetone	2 minutes	184.8 torr
CCl_4	3+ minutes	76.7 torr
alcohol	4+ minutes	43.9 torr
water	little change after 10 minutes	17.5 torr

C. Obtain two sealed tubes containing an equilibrium mixture of nitrogen dioxide (NO_2) and dinitrogen tetroxide (N_2O_4). These can be ordered from a scientific supply house, or you can prepare them (see the source cited later in this discussion). Prepare two water baths, one near the boiling point and the other near 0°C. Call attention to the fact that the two gases in the tubes at room temperature are in equilibrium:

$$2NO_2 \leftrightharpoons N_2O_4$$

Tell students that NO_2 is a red-brown gas and N_2O_4 is colorless. The color of the equilibrium mixture at room temperature indicates that there is a considerable amount of NO_2 present.

Place one of the tubes in each of the water baths. Leave them immersed until the contents have had time to respond to the temperature change. Remove the tubes from the water baths and note the colors. The tube in ice water will have become noticeably lighter in color, indicating that the equilibrium has shifted in favor of the colorless gas, N_2O_4. The tube in hot water will have become quite dark, indicating that the equilibrium has shifted in favor of the red-brown gas, NO_2. Switch the tubes from one bath to the other so students can note the results.

When both tubes are allowed to come to room temperature once again, the original equilibrium color can be observed.

This is a dramatic demonstration of the effect of temperature upon an equilibrium system.

If students are sufficiently interested to want to make their own "temperature tubes," details can be found in the *Journal of Chemical Education*, **35,** 527 (1958) in an article by R. D. Eddy and S. R. Scholes, Jr.

2. Draw the students' attention to the photograph of a food label on p. 93 of the text. See if they can guess what common food bore this label.

EXPERIMENT 3-1
LEARNING TO RECOGNIZE DYNAMIC EQUILIBRIUM

Purposes:

To introduce students to the method of recognizing a system at equilibrium; to provide a background for introducing the concept of vapor pressure; and to raise questions about the factors which may affect an equilibrium system.

Materials: (per student)

6 100-cc graduated cylinders
3 rubber bands
plastic wrap or aluminum foil
10 cc each of
  acetone
  ethyl alcohol (C_2H_5OH)
  distilled water
8 flat-bottom vials, approximately 2 cm in diameter and 4 cm tall
2 250-cc beakers
2 copper sulfate crystals
6 cc concentrated ammonia water
phenolphthalein
2 cc solid ammonium carbonate

Prelab Discussion:

Stress the importance of using the quantities of materials called for in the laboratory manual. With the recommended quantities, results are good; within a period of 35 to 40 minutes the systems have reached the point where no further changes are visible. This is important since the purpose of the experiment is to enable students to learn that when no further changes are observable, a condition of dynamic equilibrium exists in a closed system.

It is well to mention the purpose of the phenolphthalein indicator in Part III. Ammonia gas dissolves in water to produce a basic solution. Phenolphthalein turns red in the presence of a base. By having the indicator in the ammonia solution to begin with, the red color is apparent. By placing the same amount of indicator in the water, evidence of the movement of ammonia molecules from one vial to the other is easily obtained. When the color in the two vials remains constant, dynamic equilibrium has been established.

Postlab Discussion:

Points to stress:

1. Equilibrium is a dynamic process.

2. Equilibrium occurs only when a system is closed.

3. When some identifiable property, such as color, reaches a constant value and there are no further changes in the visible properties of the system, then equilibrium has been established. This is the point to the experiment: to be able to recognize whether a system is or is not in equilibrium.

Some discussion of the simultaneous processes is in order. Students should understand that the *rates* of these two processes are equal at equilibrium, but that the *amounts* or concentrations of vapor and liquid, or dissolved and undissolved solid, are not necessarily equal.

Questions:

1. If matter enters or leaves a system, then the rates of the two opposing processes involved cannot become equal. Equal rates are a necessary condition for equilibrium.
2. When no further changes occur in the observable properties of a system.
3. Part I: evaporation and condensation
   Part II: dissolving and recrystallizing
   Part III: gas escaping from solution and dissolving in water
4. The volume of the liquid decreases until the amount of vapor in the cylinder has reached the point where the rate of condensation is equal to the rate of evaporation. The rate of condensation is slow at first, when there is little vapor present, but it increases until it is equal to the constant rate of evaporation. The level of the liquid will never equal the original level because at equilibrium some of the substance remains in vapor form.
5. The safest way to answer this question is through experiment. Initially the rate of evaporation will be fast in the warm tube. However, this does not mean that equilibrium is reached sooner than in a cold tube, since, for one thing, more molecules must evaporate before equilibrium is reached.
6. Raising the temperature will increase the rate of activity at equilibrium in each system. The appearance at equilibrium and the time required could change. However, there would be no further change in appearance once equilibrium had been established.
7. No, because the test tubes were open, and to achieve an equilibrium a system must be closed.

EXPERIMENT 3-2
THE PROCESS OF BOILING

Purposes:

To provide opportunity for students to observe a familiar phenomenon under controlled conditions; to furnish experimental data which may make the boiling process more understandable; to raise questions which may be pursued in free labs.

Materials: (per student)

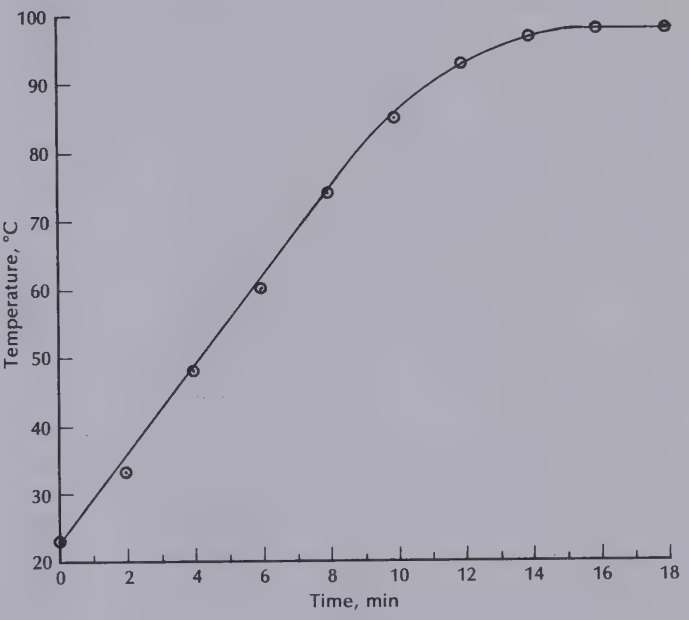
1 250-cc beaker
1 thermometer
1 stopwatch or timer (students who have watches with sweep second hands may use these)
1 burner
1 ringstand with ring
1 asbestos mat

Prelab Discussion:

Caution students about the proper handling of hot water. Remind them to observe very carefully from the beginning of the experiment. There is a tendency to overlook details because boiling water is so familiar, yet most persons never have really seen the preboiling convection currents, the variation in bubble size, etc.

Suggest that after the water has come to a full boil the flame be turned up higher. Check the temperature to see if increasing the heat causes any further rise in temperature. If the water is already boiling vigorously, the temperature will not rise. This helps to emphasize the point.

Analysis of Data:



1. Time is the independent variable, because time determines the temperature. The line is straight until it begins to level off near the boiling temperature. It does not pass through the origin since the water was not at 0 °C when the heating began.
2. The slope of the straight portion of the line will vary, depending upon the amount of water used and the rate at which heat is applied. The important point is that each student be able to interpret his own slope: for each minute of time the water is heated, the temperature rises a certain number of degrees.
3. Yes, there is a plateau. Once the vapor pressure of a liquid equals atmospheric pressure, all the energy being supplied in the form of heat is used to overcome the intermolecular forces and vaporize the liquid. During this phase change, there is no temperature change. Ask the students to compare the plateau on the water curve with the plateau on the curve for the melting of paradichlorobenzene.

Questions:

1. The first bubbles are quite small. These are composed of dissolved air, which is less soluble in warm water than in cold.
2. As the water nears the boiling point, the bubbles seen at this time are larger than the air bubbles seen earlier. These are composed of water vapor forming within the liquid water.
3. Bubbles may disappear before reaching the surface because the vapor inside cools enough to condense into the liquid state once more. The water at the bottom will tend to be hotter than that at the top until a vigorous boil is reached.
4. The vapor pressure increases with increasing temperature until it equals atmospheric pressure. At the boiling point, bubbles of water vapor can form in the interior of the liquid. (Large bubbles rising from the bottom before the top surface has started to boil indicate that the water at the bottom is already at the boiling point but the water above is cooler.)
5. Increasing the surface pressure raises the boiling temperature, because more heat energy must be supplied to raise the vapor pressure of the liquid to the point that it equals the surface pressure. Pressure cookers employ this principle.
6. It should boil at a lower temperature on top of the mountain because atmospheric pressure will be less at that altitude. The lower the pressure on the surface of the water, the lower the boiling temperature.
7. Students may need to do a free lab to investigate this question or do some reference reading. Adding solids which dissolve in water causes the boiling point of the water to increase. Suggest that they consult other books and read about Raoult's Law.
8. Boiling is the state in which the internal vapor pressure of a liquid equals that of the pressure on the surface of the liquid. It is characterized by the formation of vapor bubbles in the interior of the liquid.

Suggestions for Additional Investigation:

For your information, bubble formation occurs about a "seed." The "seed" may be impurities such as dust or grease on the glass, or irregularities in the glass surface. Take two beakers that have been thoroughly washed and dried, being careful that no lint is in either one. Make a finger print on the inside bottom of one. Pour a colorless carbonated beverage into both beakers. Bubbles form about the impurities left by the finger print. Salt sprinkled into the beverage has the same effect. Fizzing is greatly increased because there are more "seeds" about which dissolved gas can collect. Chalkdust, sawdust, and other impurities have the same effect. With a few such hints, this can make a most interesting free lab for interested students.

EXPERIMENT 3-3
THE CHEMICAL PROPERTIES OF WATER

Purposes:

To provide an opportunity for students to observe the action of water with an active metal, a metallic oxide, and a nonmetallic oxide; to provide practice in recognizing types of chemical reactions; to provide practice in writing word equations.

Materials: (per student)

1 test tube, 18 x 150 mm
1 watch glass
1 gas collecting bottle, 60 cc
1 glass plate to cover bottle
1 deflagrating spoon
1 small piece of Ca
1 small lump of CaO (FRESH!)
1 small lump of roll sulfur
a few strips of neutral litmus paper
a few wood splints
1 burner
a few crystals of $Pb(NO_3)_2$
a few crystals of KI

Prelab Discussion:

Students will have used litmus paper by now, but it might be well to ask them to review its color changes in the presence of an acid and a base. The technique for burning sulfur in a deflagrating spoon may be demonstrated. Caution them to clean the spoon by burning the sulfur completely; for this reason you may wish to insure that the small lump of sulfur is very small, indeed!

Postlab Discussion:

Discuss results of the experiment. Help student to generalize as he attempts to list the chemical properties of water. He may need assistance in completing the word equations.

Questions:

1. Part I: water + calcium →
calcium hydroxide + hydrogen
Part II: water + calcium oxide →
calcium hydroxide
Part IV: lead nitrate + potassium iodide →
lead iodide + potassium nitrate
2. Water reacts with some metals to produce hydrogen and a base. Water reacts with some metallic oxides to produce a base. Water reacts with some nonmetallic oxides to produce an acid. Water promotes some chemical changes.
3. Sweaty players could come in contact with the lime and the reaction between the lime and the water on their skin could produce a burn. This would be

especially dangerous if they got some of the lime in their eyes. One meaning of *slake* is "to become less violent." Calcium hydroxide does not readily dissolve in water, so players need not fear a reaction.

4. Unpainted iron would probably rust sooner because moisture and oxygen from the atmosphere could come in contact with it. Painting would cover the iron and thus reduce the opportunity for contact.

5. Single-acting baking powder reacts as soon as it is moistened. The baking powder in the can must have absorbed moisture from the air and reacted, releasing carbon dioxide. Therefore, no carbon dioxide was available to make the biscuits rise in the oven. The most commonly used type of baking powder is double-acting, which requires both moisture and heat in order to react; it would be more stable even if the can were left open.

EXPERIMENT 3-4
ENERGY CHANGES IN SOLUTIONS

Purposes:

To enable students to actually observe the process of dissolving and to become familiar with changes that occur during solution; to provide practice in proposing hypotheses to explain observed changes.

Materials: (per student)

30 cc distilled water
About 1 gram of each of the following solids:
  NH_4Cl
  $Ca(C_2H_3O_2)_2$
  KNO_3
  NH_4NO_3
  KOH
  $CuSO_4$ (anhydrous)
6 100 x 13 mm test tubes
1 thermometer
6 glass slides
1 glass stirring rod or eye dropper
1 magnifying glass or microscope

Prelab Discussion:

Emphasize that in step 3, only a few crystals of the solids are to be placed in the water (and only 1 pellet of KOH). The purpose of this step is to *observe* the dissolving process. Review the evidences for chemical change (Experiment 2-2): evolution of a gas, liberation of heat, formation of a new substance.

Stress the importance of recording the *maximum* temperature change as solids dissolve in step 5.

Caution the students that the stirring rod or eye dropper should be rinsed clean each time it is used to place a drop of solution on a glass slide.

Results:

Student results will vary, but the following table gives you an idea of the kinds of results that may be expected.

Solid	*Temperature Change Observed*
NH_4Cl	22.9°C to 17.5°C
KNO_3	22.0°C to 17.0°C
KOH	27.9°C to 47.9°C
$CuSO_4$	26.0°C to 29.8°C

Questions:

1. Solids causing temperature to increase: potassium hydroxide and anhydrous copper sulfate.

Solids causing temperature to decrease: ammonium chloride, calcium acetate, potassium nitrate, and ammonium nitrate.

2. Solids that dissolve by decreasing the water temperature would be more soluble in hot water. Heat is required to dissolve these solids, so increasing the temperature of the water would provide more energy and the solution process would be favored.

3. Referring to the three steps in the solution mechanism in the textbook, endothermically dissolving solids absorb more energy in step one than they release in step three. Such solids have strong bonds holding the particles in the crystals. Exothermically dissolving solids absorb less energy in step one than they release in step three. Such solids have weak bonds holding the particles in the crystal.

4. There is no apparent difference in the structure of a solid after it has been dissolved and recrystallized. Students may be misled at this point, because prior to being dissolved the crystals may be held together in clusters, whereas after recrystallization more separate and individual crystals may be evident. However, the submicroscopic structure has not been changed. For most solids, dissolving is a phase (physical) change in that the composition of the matter is not altered.

EXPERIMENT 3-5
UNSATURATED, SATURATED, AND SUPERSATURATED SOLUTIONS

Purposes:

To observe the three types of solutions with particular emphasis on the instability of the supersaturated solution; to provide experience in observing that a saturated solution is one in which the dissolved solute exists in equilibrium with the undissolved solute; to provide qualitative evidence for energy changes during the dissolving and crystallizing processes.

Materials: (per student)

1 test tube, 18 x 150 mm
1 beaker, 250 cc
1 burner
1 test tube holder
1 balance
15 g of sodium thiosulfate pentahydrate, $Na_2S_2O_3 \cdot 5H_2O$

Prelab Discussion:

Stress the importance of careful weighing of the solid and precise measurement of water as both are critical in this experiment. Also caution students that the mixture must be heated only enough to aid the solution process. If the solution is overheated, water will be lost and the results will not be good.

Students must *not* use thermometers in any part of this experiment. All temperatures should be measured qualitatively by touching the bottom of the test tube to the inner wrist.

Postlab Discussion:

This is one of the most fascinating experiments for students. The fact that such a large amount of solid will dissolve in such a small volume of water is intriguing to them. Also, the crystal formations that appear when a small seed crystal is placed in the supersaturated solution are most interesting, as is the heat that is evolved during this process.

The points to stress are (1) that a saturated solution is one in which there is a dynamic equilibrium between the dissolving and the crystallizing process and (2) that supersaturated solutions are highly unstable. Frequently, just tapping the test tube containing the supersaturated solution causes the excess solute to crystallize.

Questions:

1. First, sodium thiosulfate must have the type of bonds that make it quite soluble in water. The crystals are ionic, which supports this hypothesis. Second, the crystals contain 5 molecules of water of hydration, which are held in loose union with the crystal. This adds to the total volume of water. The dot between the two parts of the formula should cause students to raise questions. Suggest that student consult resource materials to learn more about such compounds. For a clue, if one is needed, tell them the term *hydrate* may give them a lead into the subject.

2. If students have observed carefully, they will have noted that the water temperature was lowered. The overall endothermic effect is of course related to the great attraction the ions in the crystal have for each other; energy must be absorbed to move them farther apart. This energy comes from the water, which cools. Energy is released by hydration of the dissolving ions, but the amount of this energy is less than that required to pull the ions apart.

3. Crystallization releases heat. As unlike particles (ions) move closer together, the potential energy of the system is lowered. The excess energy that was absorbed to move the ions apart is now released as they assume the more stable position in the crystal. Remind students of the opposing tendencies: minimum energy versus maximum randomness.

4. The ions are much more randomly distributed when the solute is dissolved than when the crystal is whole. As the crystal dissolves, the ions become dispersed and there are many more ways in which they can be arranged than when they are in the orderly arrangement of the crystal.

5. The saturated solution is most stable because the crystal and the dissolved solute are in equilibrium. The supersaturated solution is most unstable. This is evident from the fact that sometimes the excess solute falls out during the cooling process.

6. The dissolved solute contains more potential energy than the crystal, since it takes energy to separate the ions of the crystal.

CHAPTER 4

ACIDS, BASES, AND SALTS

Goals:

1. To introduce the ionization theory as our best explanation for the behavior of acids, bases, and salts.
2. To apply the concept of equilibrium to solutions of electrolytes.
3. To provide laboratory experience with the properties of acids, bases, and salts.
4. To introduce buffers and to relate these solutions to their importance in the human body.
5. To introduce the use of chemical indicators in determining the pH of solutions.

Teaching Hints:

1. Continue to check with students who are not working up to capacity.

2. Emphasize the importance of doing some creative work in connection with each topic. We especially want to encourage open labs, for it is in these activities that sutdents gain the most experience in designing experiments to test hypotheses they themselves have proposed. They learn how to ask the proper questions and to solve problems using an orderly process.

3. Relate class experiences to daily living at every possible opportunity.

4. The use of Styrofoam models is very helpful in teaching the concept of proton donors and proton

acceptors. See the Appendices of this Guide for directions on making hydrogen, oxygen, water, and ammonia molecules.

6. Students have not yet learned how to write chemical equations. However, it was our experience during the pilot use of this course that this is not a problem. From the beginning we have used word equations. Students seem to accept these, and as we begin to use chemical equations the transition is not difficult. At this point we do not expect students to write chemical equations by themselves, and they seem to understand them as presented in the textbook.

7. It is recommended that students memorize the names and formulas for the commonly used acids and bases. However, some teachers prefer to permit the use of the tables in the text and do not require memorization. This is left to your personal preference.

EXPERIMENT 4-1
ELECTRICAL CONDUCTIVITY OF LIQUIDS

Purposes:

To demonstrate that liquids differ in their ability to conduct an electric current; to show that factors affecting conductivity of solutions include the nature of the solvent and the nature of the solute; to provide a background for the discussion of the Arrhenius concept of ionization.

Recommendations:

You may wish to do this experiment as a demonstration. However, it is preferable for each student to do it for himself if safety can be assured. It will depend upon the space available in your laboratory and whether the students are moving at their own rate. If it is done as a demonstration for the whole class, the slower paced students can take the data and file it for later analysis.

It is suggested that the conductivity be measured quantitatively for two reasons: (1) this must be done if the slight conductivity of "pure" water is to be demonstrated; (2) it provides evidence that even strong electrolytes vary in their ability to conduct current. The experiment may be done qualitatively with a light bulb powered by a 110-volt power supply or by a battery, but it is very difficult to distinguish between good and fair electrolytes, because a weak current can flow through the bulb without lighting it. However, this method is better than no demonstration. Do the best you can with the equipment available.

The apparatus depicted in the illustration in the laboratory manual can be used as a basis for an individual kit. The multimeter (volt-ohm- milliammeter) has a built-in power source. Wide-mouth jars with covers can be the test containers and can also serve for storage of the solutions and solvents.

One caution in the use of a multimeter is to turn the setting to some position other than a resistance setting for storage. Also, remember that resistance is the reciprocal of conductivity — the higher the resistance reading, the smaller the conductivity.

Materials:

24 beakers, 250 cc, *or* wide-mouth jars with covers
1 conductivity apparatus
hook-up wires
1 multimeter (volt-ohm-milliammeter)
200 cc of each of the following:
  distilled water
  tap water
  distilled water with a few crystals of NaCl dissolved in it
  0.1M NcCl (6 g/l)
  0.1M KNO_3 (10 g/l)
  0.1M $CuSO_4$ (25.0 g pentahydrate/l)
  0.1M HCl (8.5 cc conc./l)
  0.1M H_2SO_4 (5.6 cc conc./l)
  0.1M HNO_3 (6.4 cc conc./l)
  0.1M CH_3COOH (5.7 cc glacial/l)
  0.1M NaOH (4 g/l)
  0.1M KOH (5.6 g/l)
  0.1M ammonia water (6.6 cc conc./l)
  $Ca(OH)_2$—saturated
  ethanol
  CCl_4
  benzene
  acetone
  glacial acetic acid
  sugar + water
  sugar + ethanol
  NaCl + ethanol
  HCl + benzene
  glacial acetic acid, with water added as students watch current change

Expected Results:

The actual readings you get will depend upon the equipment used. Also, distilled water of the kind typically available will show a higher conductivity than an absolutely pure sample of water because it has traces of impurities in it. So it is well to point out that pure water has less conductivity than a typical sample of distilled water displays.

Questions:

1. Acids were the strongest electrolytes, salts the next, and bases the weakest. Ethanol, CCl_4, benzene, and glacial acetic acid were non-electrolytes.

2. The concentrations of the salt, acid, and base solutions were the same. The concentration is related to the total possible number of ions, if ionization is complete.

3. Glacial acetic acid is 99+% pure and does not ionize. As polar water molecules are added, bonds break in some acetic acid molecules and ions form. This implies that glacial acetic acid is also a polar substance. If the hypothesis is correct, it should apply to other substances that dissolve in water but are nonconductors in pure liquid form. Unfortunately not very many such substances exist. Most organic acids are solids at room temperature. Anhydrous citric acid can be melted (153°C) and then tested for conductivity. Oleic acid is another substance that could be tested. Tartaric acid can also be melted (170°C) and tested. Encourage students to search for other possibilities in the tables of a handbook.

4. Good conductors contain many ions (charged particles). Poor conductors contain a low concentration of ions. Nonconductors contain no ions.

EXPERIMENT 4-2
PREPARATION OF A SALT FROM A TYPICAL ACID AND BASE

Purposes:

To provide laboratory experience in studying the properties of a typical acid (HCl) and a typical base (NaOH); to combine these in proper proportion to obtain a typical salt (NaCl); to provide evidence that properties of compounds change when a chemical reaction occurs; to introduce a general method for the preparation of salts.

Materials: (per student)

40 to 50 cc 3M HCl (256.5 cc concentrated reagent per liter of solution)
40 to 50 cc 3M NaOH (120 g per liter of solution)
0.2 g $NaHCO_3$ (baking soda)
phenolphthalein indicator solution (1 g in 100 cc 50% ethanol)
neutral litmus paper
1 or 2 small chips $CaCO_3$ (marble chips)
6 100 x 13 mm test tubes
1 evaporating dish
1 graduated cylinder
1 stirring rod
1 hand lens
1 glass slide

Prelab Discussion:

Caution students about handling acids and bases properly. Remind students that in Part III, step 2, they may need to add just a bit more than 9 cc of HCl. The important point is to add just enough HCl to the base so that the solution becomes neutral. Also suggest that they review the technique for evaporating a liquid in an evaporating dish.

Questions:

1. Acids turn litmus red; do not change the color of phenolphthalein; react with bases to form salts; and react with carbonates to produce carbon dioxide, another salt, and water.

2. Bases turn litmus blue; change the color of phenolphthalein to red; react with acids to form salts.

3. Hydrochloric acid + sodium hydroxide →
sodium chloride + water

4. The first element named in the salt comes from the base and the second element named comes from the acid.

a. To produce potassium chloride:
hydrochloric acid + potassium hydroxide →
potassium chloride + water

b. To produce sodium sulfate:
sulfuric acid + sodium hydroxide →
sodium sulfate + water

c. To produce potassium nitrate:
nitric acid + potassium hydroxide →
potassium nitrate + water

EXPERIMENT 4-3
ACID-BASE INDICATORS

Purposes:

To introduce different acid-base indicators; to provide experience in determining the pH of solutions by indicator colors.

Materials: (per student)

dropping-bottle quantities of the following:

0.1M HCl (pH 1)	0.1M NaOH (pH 13)
0.01M HCl (pH 2)	0.01M NaOH (pH 12)
0.001M HCl (pH 3)	0.001M NaOH (pH 11)
0.0001M HCl (pH 4)	0.0001M NaOH (pH 10)
distilled water (pH 7)	

10 cc phenolphthalein (1 g in 100 cc 50% ethanol)
10 cc methyl orange (0.1 g in 100 cc water)
10 cc universal indicator (Gramercy) with color chart (If this is not available, use litmus solution. Litmus cannot make as fine distinctions.)
1 roll Hydrion paper with color chart
3 spot plates *or* piece of clear plastic plus grease pencil

unknowns, if desired:
  0.1M acetic acid solution, buffered with sodium acetate — pH 4.5
    (1.36 g sodium acetate + 10 ml 1.0M acetic acid per 100 cc of solution)
  0.05M sodium carbonate solution — pH 11.5
    (0.62 g $Na_2CO_3 \cdot H_2O$ per 100 cc of solution)

Prelab Discussion:

Instruct students to test standards with Hydrion paper prior to addition of other indicators. If time permits, students can determine the pH of the "unknowns" by comparison with the standard colors on the spot plates.

Postlab Discussion:

Let students report results for comparison purposes. Results are usually quite good, if original solutions were prepared carefully. The Hydrion paper test is probably the most useful for the general student. The degree to which you go into the technical definition of pH will depend upon the students. For most classes, the general definition of degree of acidity or basicity on the 1-to-14 scale is sufficient for understanding of the experiments in a general course. Do not require the use of logarithms unless your students are good in math and are already familiar with their use.

Questions:

1. The more dilute the acid, the higher the pH number (limiting value 7). The more dilute a base, the lower the pH number (limiting value 7)

2. The following are the pH values if the acids ionize completely.
   a. 5; b. 2; c. 0

3. Help students to generalize from the pH values marked on the dropping bottles to the rule that for strong bases the pH equals the number of decimal places in the molarity subtracted from 14.
   a. $14 - 1 = 13$; b. $14 - 0 = 14$;
   c. $14 - 4 = 10$

4. This will depend somewhat upon the quality of the students' results. Most will probably feel that the Hydrion paper is best.

5. Litmus is made from lichens (primitive plants composed of combined algae and fungi); phenolphthalein is a coal tar product; methyl orange is the sodium salt of helianthine.

6. You don't need all the dilutions to see the general pattern of indicator colors; it is impossible to make up "all possible" dilutions as there are an infinite number of them.

EXPERIMENT 4-4
HYDROLYSIS, THE COMMON ION, AND BUFFERS

Purposes:

To provide experimental evidence for the existence of salts that dissolve to form acidic or basic solutions as well as those that form neutral solutions; to demonstrate the effect on the pH of adding a common ion to a solution of a weak acid; to demonstrate the role of buffers in controlling the pH of solutions.

Materials: (per student)

2 cc of each of the following 1.0M solutions:
  sodium chloride, $NaCl$ (11.7 g/200 cc)
  sodium sulfate anhydrous, Na_2SO_4 (28.4 g/200 cc)
    or crystalline $Na_2SO_4 \cdot 10H_2O$, (64.4 g/200 c)
  potassium nitrate, KNO_3 (20 g/200 cc)
  sodium acetate, $NaC_2H_3O_2 \cdot 3H_2O$ (28 g/200 cc)
  sodium carbonate, $Na_2CO_3 \cdot 10H_2O$ (58 g/200 cc)
  sodium phosphate, $Na_3PO_4 \cdot 10H_2O$ (69 g/200 cc)
  ammonium chloride, NH_4Cl (12 g/200 cc)
  aluminum sulfate, $Al_2(SO_4)_3 \cdot 18H_2O$ (136 g/200 cc)
  copper sulfate, $CuSO_4 \cdot 5H_2O$ (50 g/200 cc)
1 glass slide or watch glass
9 13 x 100 mm test tubes
10 18 x 150 mm test tubes
methyl orange (0.1 g in 100 cc of water)
phenolphthalein (1.0 g in 100 cc 50% ethanol)
1 gram, or ¼ teaspoon, sodium acetate
1 gram, or ¼ teaspoon, ammonium chloride
1 gram, or ¼ teaspoon, potassium chloride
20 cc 1M acetic acid (12 cc glacial acetic acid/200 cc solution)
20 cc 1M ammonia water (7 cc conc. solution/100 cc)
20 cc 1M HCl (90 cc concentrated HCl/liter solution)
10 cc distilled water
10 cc buffer solution, pH 7
    (Note: It is preferable to purchase buffer solution of pH 7, but you can make your own by adding 148.15 cc 0.1M NaOH to 250.0 cc 0.1M KH_2PO_4 (3.4 g/250 cc solution) and diluting to make 500.0 cc solution.)
10 cc 0.1M NaOH (4 g/liter)
10 cc 0.1M HCl (8.6 cc conc. acid/liter)

Prelab Discussion:

Students should be familiar with the terms introduced in the laboratory manual. The influence of the change in H_3O^+ and OH^- concentration on equilibrium should be mentioned and students should be encouraged to think about this as they perform the experiment. The importance of buffers in living systems is a topic which should be explored through both pre-and postlab class discussion of the experiment. Outside reading on this topic is highly recommended.

Postlab Discussion:

Students should be led to see that a salt consists of a positive ion from a base and a negative ion from an acid. When the salt hydrolyzes in water, some of both the parent acid and the parent base forms. If the acid is strong and the base weak, the salt solution is acidic; if the acid is weak and the base strong, the salt solution is basic; if both the acid and the base are either weak or strong, the salt solution is neutral (for all practical purposes). The latter is not always true if very sensitive tests are run.

Common Ion (Part II):
When sodium acetate is added to a solution of acetic acid, the pH goes up. The concentration of undissociated acetic acid is increased; or the degree of dissociation decreases. The addition of acetate ions enables more of the H_3O^+ ions to be tied up, thus decreasing the acidity of the solution. Equilibrium shifts to favor molecules, not ions.

No effect is noticed when Cl^- ions are added to a dilute aqueous solution of HCl. This is because the HCl is considered to be completely ionized and no equilibrium exists between HCl and H_3O^+ and Cl^- ions. Thus, addition of a common ion does not appreciably affect the concentration of H_3O^+ nor the pH.

The addition of NH_4Cl to ammonia water decreases pH. The NH_4^+ ion is the common ion and shifts the equilibrium to the left, thus reducing the concentration of OH^- ions and lowering the pH.

$$NH_3 + HOH \leftrightharpoons NH_4^+ + OH^-$$

Buffers (Part III):
The experiment in the lab manual is a dramatic example of the effect of buffers on pH. Orange is the intermediate color for methyl orange. The indicator becomes pink at a pH of 3.1 and yellow at a pH of 4.4. The number of drops varies with student data, but usually one drop of 0.1M HCl causes the pink color to show in the distilled water solution and from 12 to 18 drops are required to cause the buffer to show a pink color!

The color change in methyl orange is difficult for some students to detect. For some types of color blindness this is impossible. Usually these students can see the color change when bromthymol blue indicator is used.

Questions:

1. See first paragraph of postlab discussion.
2. $Na_2CO_3 + 2H_2O \rightarrow H_2CO_3 + 2\ Na^+ + 2\ OH^-$
   $NH_4Cl + H_2O \rightarrow NH_3 + H_3O^+ + Cl^-$
3. a. Once an equilibrium has been established between acetic acid molecules and acetate ions, the addition of more acetate ions places a stress upon the system. To relieve this stress, according to Le Chatelier's Principle, acetate ions unite with the proton from hydronium ions to form more acetic acid molecules. Acetate ions are good bases, and take protons readily from the hydronium ions. This reduces the number of hydronium ions in the solution, making it less acidic.

   b. There is no equilibrium between the hydrogen chloride molecules and the hydronium ions. Hydrogen chloride is totally ionized in water solution. Since there is no equilibrium existing in the first place, there is no change in pH when chloride ions are added, because hydronium ions are neither increased or decreased in number.
4. See example in the laboratory manual.

CHAPTER 5

THE DEVELOPMENT OF ATOMIC THEORY

Goals:

1. To trace the historical development of the modern atomic theory.
2. To provide an understanding of the importance of individual effort in the progress of science.
3. To point up the role of imagination and creativity in the formulation of scientific theories.
4. To provide laboratory experiences that will make theory more meaningful.

Teaching Hints:

1. Some students who do not enjoy the study of history may fail to see the importance of this presentation of the atomic theory. Try to lead them to understand that the modern achievements in science and technology rest upon discoveries made in the past. Encourage students to read the biographies of several of the scientists mentioned in the chapter. Jaffe's accounts in *Crucibles: The Story of Chemistry* (see chapter bibliography) are the most fascinating.

   If students can realize that scientists are people and that trial and error have played a large part in building the structure of science as we know it today, perhaps they will see that science is a human activity.
2. Set up a demonstration on the electrical nature of matter. Include rubber rods, glass rods, fur, a balloon, and an electroscope. Students should observe this demonstration prior to studying Sec. 5-5.

3. If possible set up a demonstration so students can observe the spectrum of hydrogen. Spectrum tubes of hydrogen, helium, mercury, and others, along with the proper power supply, are available from many scientific supply houses.

This demonstration is helpful if done when students begin their study of the Bohr atom (Secs. 5-11 and 5-12).

4. Continue to encourage creative work.

EXPERIMENT 5-1
BUILDING A MODEL FROM INDIRECT EVIDENCE

Purposes:

To provide experience in obtaining evidence by indirect methods; to illustrate the volume of information that can be obtained by indirect methods; to emphasize the importance of imagination in science; to provide an opportunity for students to construct a model so that they may appreciate the use of scientific models in describing abstract ideas.

Materials: (per class)

A sufficient number of "black boxes" should be available so that each pair of students can work on one simultaneously. These may be made up from a variety of empty cardboard boxes. Sizes should vary as well as the contents and internal patterns.

Cardboard partitions may be held in place by the use of masking tape. Some boxes should contain no partitions; some should have a simple pattern; a few should be rather complex, as a maze.

Objects in the boxes should be commonly known to students. Suggestions: thumb tack, key chain, cork stopper, small magnet, small ball bearing, cotton ball, paper clip, Styrofoam sphere, wood blocks of different shapes and sizes, short pencil, empty lipstick case, etc.

It is recommended that the boxes be numbered for easy identification.

Note: Some of the pilot teachers asked a colleague to secretly seal the objects in the boxes so that the pilot teacher approached the problem in the same way that the student approached it. They found the experience stimulating!

Prelab Discussion:

Make it clear to students that the objects are commonly known to them. Suggest that more may be learned about the contents of the box by careful, controlled movements than by violent shaking. If possible, have some empty boxes identical to some of the sealed boxes which students may use as a control for comparison purposes.

Balances, metric rules, and magnets should be available for use if students ask for them.

Results and Postlab Discussion:

Teachers vary in their opinions about whether or not students should be allowed to learn what is actually in the boxes. Experience has indicated that students are much more likely to have confidence in their ability to obtain indirect evidence if they are eventually permitted to open the boxes. It is surprising how accurate some of the predictions can be, even to determining the size and shape of objects.

After students have submitted reports on the boxes, one or two boxes might be opened in each class. It is suggested that the predictions for a given box be listed on the board prior to opening the box. This makes for excitement and suspense and makes comparison with prediction and actual contents easier.

Students should be helped to realize that both imagination and experience enter into their predictions. They tend to associate certain objects with certain sounds. Likewise, they associate a certain color with wood or metal.

Point out that scientists use similar methods in arriving at the models for abstract concepts and that many such models will be studied as we trace the development of the modern theory for atomic structure.

Questions:

1. Because the contents are unknown and must be identified by indirect observations.

2. Any unsolved problems: cure for cancer, cause of colds, etc.

3. This question does not have a certain answer. Some students may suggest color as one property that cannot be investigated indirectly, although it can be observed indirectly by using instruments sensitive to radiant energy. Matter may have properties of which we are unaware and will be unaware until someone proposes the existence of the property and also suggests instruments for observing it indirectly.

Sometimes the observational instrument comes first and then all kinds of ways to use it for indirect observation are invented. The electron microscope is a prime example.

EXPERIMENT 5-2
IDENTIFICATION OF IONS BY COLOR

Purposes:

To introduce the techniques for performing flame tests, cobalt nitrate tests, and borax bead tests; to provide direct evidence that elements can be identified by the color they impart.

Materials: (per student)

1 platinum wire, approx. 3 cm long, in a glass tube holder, and with small loop at the end
6 to 12 test tubes (100 x 13 mm)
1 small vial of powdered borax
1 charcoal block
1 blowpipe
5 cc of each of the following solutions (all 0.1M):

Solution	*Formula of solid to be dissolved*	*Grams/liter to make 0.1M*
$NaNO_3$	$NaNO_3$	8.5
KNO_3	KNO_3	10.1
$Ca(NO_3)_2$	$Ca(NO_3)_2 \cdot 4H_2O$	23.6
$Ba(NO_3)_2$	$Ba(NO_3)_2$	26.1
$Sr(NO_3)_2$	$Sr(NO_3)_2$	18.7
$Cu(NO_3)_2$	$Cu(NO_3)_2 \cdot 3H_2O$	24.0
$Mg(NO_3)_2$	$Mg(NO_3)_2 \cdot 6H_2O$	25.6
$Co(NO_3)_2$	$Co(NO_3)_2 \cdot 6H_2O$	29.0
$Cr(NO_3)_3$	$Cr(NO_3)_3 \cdot 9H_2O$	40.0
$Mn(NO_3)_2$	$Mn(NO_3)_2$	17.9
$Ni(NO_3)_2$	$Ni(NO_3)_2 \cdot 6H_2O$	29.0
$Al(NO_3)_3$	$Al(NO_3)_3 \cdot 9H_2O$	37.5
$Zn(NO_3)_2$	$Zn(NO_3)_2 \cdot 6H_2O$	30.0

Prelab Discussion:

Demonstrate the proper techniques for performing each of the tests used in this experiment. Emphasize the importance of cleaning the platinum wire between tests and the necessity for scraping prior materials from the charcoal block.

Demonstration:

It is highly desirable that students view the hydrogen spectrum through a spectroscope or with a diffraction grating, if possible. The spectroscope or grating, hydrogen tube, and spectrum tube power supply can be set up so that students may view the spectrum individually during the laboratory period. It is well to have a colored picture of the hydrogen spectrum at the spectroscope for comparison with the actual spectrum. (See Text, p. 124, top band.)

Results and Postlab Discussion:

Part I: Flame Tests

Metallic ion	*Flame color*
sodium	persistent yellow
potassium	violet, lilac
strontium	red
copper (II)	bright green
barium	yellowish green
calcium	orange to red

Emphasize the fact that upon heating, electrons are excited to a higher energy level in the atom. As they fall back to lower energy levels, they lose energy; the energy is emitted in the form of visible light which we see as the characteristic flame color of that element. Some elements emit light outside the visible range.

Part II: Borax Bead Tests

Metallic ion	*Bead color*
cobalt (II)	blue
chromium (III)	green
manganese (II)	violet
nickel (II)	brown

The chemical reactions involved in these tests and the composition of the colored compounds formed are not well understood at present.

Part III: Cobalt Nitrate Tests

Metallic ion	*Color*
aluminum	blue
zinc	green
magnesium	pink

These tests are based on the fact that when strongly heated, the cobalt nitrate decomposes to the oxide. This oxide combines with the oxides of aluminum, zinc, and magnesium to form colored complexes.

Questions:

1. Since all the compounds tested have a common negative ion (nitrate), the different colors must be due to the positive ions. This logic can be confirmed in a free lab investigation using pairs of compounds with the same positive ion but different negative ions.
2. Students should have learned by now that dissolving is a phase change and normally does not change the composition of the solute. Therefore, they would guess that the dry salts should impart the same color as the solution of that salt. Have samples of the dry salts available for testing.

Suggestions For Additional Activities:

1. Have 5 cc samples of the compounds used in this experiment in small test tubes labelled with a code letter so that students may identify at least one "unknown."
2. Be sure students do the fireworks activity as library research only! Actual experimentation is considered too dangerous for students with limited information about the composition of the fireworks.

EXPERIMENT 5-3
PAPER CHROMATOGRAPHY

Purposes:

To introduce another technique for the identification of elements; to provide an interesting laboratory experience.

Materials: (per student)

3 Petri dishes
6 to 8 filter paper discs
black ink
  (Note: Washable ink is preferable to permanent ink. Do *not* use India ink.)
universal indicator solution
1 metric ruler
25 cc of each of the following solvents:
  water
  ethyl alcohol
  ammonia water (Conc. ammonia water diluted to 4 times its initial volume)

Prelab Discussion:

Demonstrate the proper technique for cutting the tongue in the filter paper. It is very important that the tongue be just barely long enough to touch the solvent; otherwise the paper will collapse into the solvent as the process goes on. Emphasize the importance of using a very tiny drop of the ink or other liquid to be tested.

Questions:

1. Answers vary, depending upon what brand of ink is tested. However, the solvent that gives the greatest Rf values is the one in which the pigments are most soluble.
2. The chromatograms do vary with the type of solvent. This is because some pigments are more soluble in one solvent than in another.
3. The top of the ink spot is the point from which all pigments start.

Suggestions For Additional Activities:

These can be encouraged as open labs. Many different results are obtained when different substances are used.

CHAPTER 6
THE ELECTRONIC STRUCTURE OF ATOMS AND THE PERIODIC LAW

Goals:

1. To provide an elementary understanding of the theory of atomic structure as viewed at the present time.
2. To provide practice in writing electron configurations, orbital notation, and electron dot notation.
3. To lead students into an understanding of the modern periodic table of elements by having them construct their own periodic table through the study of several properties of the first 54 elements.
4. To introduce the general trends in properties of elements, either in the same period or in the same family.

Teaching Hints:

1. If students acquire a thorough understanding of this chapter, their work in the remainder of the course will be greatly simplified. Learning how to use the periodic table for interpreting the behavior of elements in chemical reactions is one of the most powerful tools in chemistry.

We have found through experience that when students do Experiments 6-1 through 6-4 as instructed in the laboratory manual, they emerge with a fundamental understanding of the periodic table that is invaluable to them. We strongly suggest that no periodic table be displayed in the room until after most students have completed these exercises.

Some students who have had a course in physical science will know that they are working on the development of the periodic table. The way you handle this will depend upon your estimate of the level of understanding of such students. However, we have found that if the teacher is frank with the student and encourages him to do the exercises according to directions, he will usually cooperate. Explain to these students that if they perform these four experiments as instructed, they will benefit because of the increased understanding gained by "doing it our way."

2. Insist that these four experiments be done *in class*. This further insures that students are not using a periodic table to guide them. Each experiment should be checked with the student as he completes it. This is important because each succeeding experience builds upon the preceding one.

3. Experiments 6-5 and 6-6 introduce techniques that will be used when the students study Chapter 11, which deals with elementary qualitative analysis. These are placed in the course at this point to provide

some laboratory work which is meaningful and at the same time provides a respite from their study of theory.

4. Students who are way behind may be allowed to skip Chapters 6 through 9 in order to complete 10 and 11 by the end of the year. These students should perform Experiments 6-5 and 6-6, which are necessary for Chapter 11.

EXPERIMENT 6-1
GROUPING ELEMENTS INTO FAMILIES ON THE BASIS OF SIMILAR ELECTRON PATTERNS

Purposes:

To provide practice in writing electron configurations; to emphasize the repeating pattern by which electron levels fill; to point up the fact that no two elements have exactly the same electron configuration; to begin a series of activities in which the student will eventually construct a periodic table.

Prelab Discussion:

Students must know how to write electron configurations prior to this activity. They must know the order for filling the levels and the orbitals and the maximum number of electrons permitted in each level and sublevel.

After students write down the filled or occupied orbitals, they should rearrange them so that all electrons in a given energy level are grouped together. For example, even though the 4*s* orbitals fill before the 3*d* orbitals, 3*d* should be written first, right next to 3*p*. This is necessary because students will be interested in the number of electrons in the *outermost* level as they group the elements.

It is strongly recommended that this exercise be done under your supervision, without access to a periodic table. Some students who have had physical science will probably be aware of what they are doing, but even these will develop a much deeper understanding of the periodic table if they develop it now as planned in this series of activities.

Results and Postlab Discussion:

Tabulations are given below for procedural steps 2 and 3:

2. a. $_{2}$He $_{10}$Ne $_{18}$Ar $_{36}$Kr $_{54}$Xe
b. $_{1}$H $_{3}$Li $_{11}$Na $_{19}$K $_{37}$Rb
c. $_{4}$Be $_{12}$Mg $_{20}$Ca $_{21}$Sc $_{22}$Ti $_{23}$V $_{24}$Cr $_{25}$Mn $_{26}$Fe $_{27}$Co $_{28}$Ni $_{29}$Cu $_{30}$Zn $_{38}$Sr $_{39}$Y $_{40}$Zr $_{41}$Nb $_{42}$Mo $_{43}$Tc $_{44}$Ru $_{45}$Rh $_{46}$Pd $_{47}$Ag $_{48}$Cd
d. $_{5}$B $_{13}$Al $_{31}$Ga $_{49}$In
e. $_{6}$C $_{14}$Si $_{32}$Ge $_{50}$Sn
f. $_{7}$N $_{15}$P $_{33}$As $_{51}$Sb
g. $_{8}$O $_{16}$S $_{34}$Se $_{52}$Te
h. $_{9}$F $_{17}$Cl $_{35}$Br $_{53}$I

3. first row: $_{21}$Sc $_{22}$Ti $_{23}$V $_{24}$Cr $_{25}$Mn $_{26}$Fe $_{27}$Co $_{28}$Ni $_{29}$Cu $_{30}$Zn
second row: $_{39}$Y $_{40}$Zr $_{41}$Nb $_{42}$Mo $_{43}$Tc $_{44}$Ru $_{45}$Rh $_{46}$Pd $_{47}$Ag $_{48}$Cd

Questions:

1. Five in a and b; four in all the rest.

2. There are either 8 or 18 elements between the ones listed in vertical columns. The significance of this rests in the rule for the order of filling levels and orbitals and in the maximum number of electrons permitted in an outermost level. This is of key importance and some time should be spent in discussing this pattern in relation to the textbook material on order of filling levels.

3. Ten. This is because *d* orbitals are filling and the maximum number of electrons permitted in a *d* orbital is 10.

4. $_{55}$Cs : $1s^2$ / $2s^2\ 2p^6$ / $3s^2\ 3p^6\ 3d^{10}$ / $4s^2\ 4p^6\ 4d^{10}$ / $5s^2\ 5p^6$ / $6s^1$

$_{57}$La: Same as Cs except outermost energy level is $6s^2$ and there should be one 4*f* orbital. In reality there is one 5*d* orbital instead.

$_{82}$Pb: Same as $_{55}$Cs except 4th level contains $4f^{14}$, 5th level contains $5d^{10}$, and 6th level contains $6s^2\ 6p^2$

$_{86}$Rn: Same as $_{82}$Pb except 6th level contains $6p^6$ instead of $6p^2$.

$_{55}$Cs would go in the second column, just under $_{37}$Rb, because it contains one electron in the outermost energy level. $_{57}$La would be placed in a position to start another horizontal row, because theoretically it begins the filling of the 4*f* orbitals. It does not fit in the present chart. $_{82}$Pb would be placed in the vertical column of elements having 4 outer shell electrons; it would go under $_{50}$Sn. Finally, $_{86}$Rn would go in the first column, under $_{54}$Xe, because the outermost shell contains a maximum number of electrons.

EXPERIMENT 6-2
THE RELATIONSHIP BETWEEN ATOMIC NUMBER AND THE RADIUS OF ATOMS

Purposes:

To introduce the students to the effect of increasing atomic number on the radius of atoms; to provide further information which will enable students to construct a periodic table based upon evidence they have assembled.

Prelab Discussion:

Be sure students know that atomic number is to be plotted on the horizontal axis. After the points are plotted, they may be joined by a series of straight lines (as in "connect-the-dots" puzzles). The reason for this is that there is no meaning to non-integer atomic numbers, so we are not interested in any atomic radius values in between the plotted points.

This activity can be assigned as homework if you wish. The effect of adding an outer energy level greatly increases the radius of an atom, and the effect of adding electrons to the same energy level results in a gradual decrease in the radius. These trends will be "discovered" by the student as he examines the finished graph.

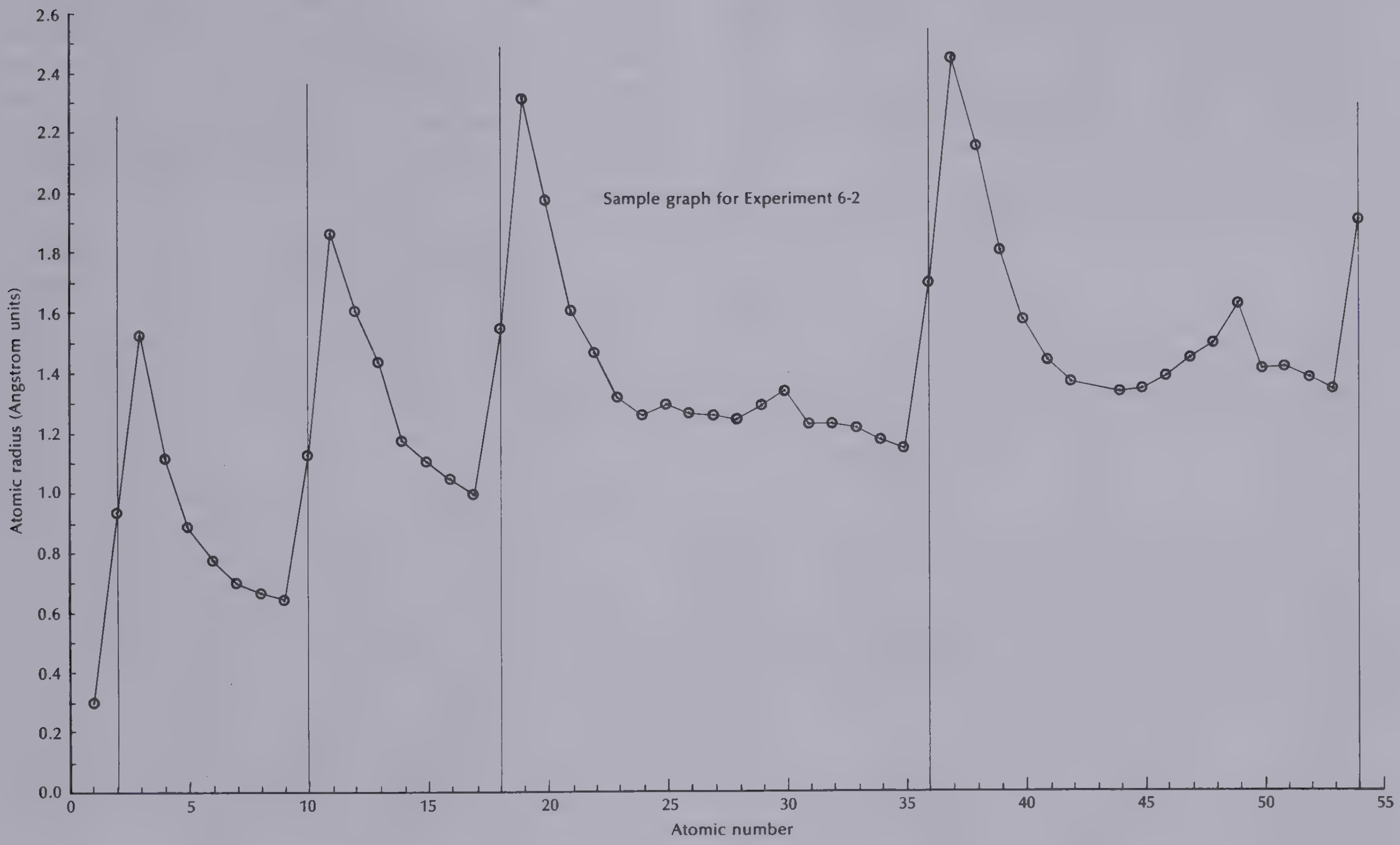
Postlab Discussion:

As the trends are studied and discussed it is helpful to have the electron configurations written in Experiment 6-1 handy for reference. The points to be stressed include:

1. Addition of a new energy level has a great effect upon the size of the atom, because the new level is much farther from the nucleus than the preceeding level;
2. Addition of electrons to orbitals within the same energy level results in a decrease in atomic size in most cases. This is because the orbitals within an energy level are fairly close together and the increasing atomic number raises the effective nuclear charge sufficiently to pull the electrons closer to the nucleus.

These concepts are so important that considerable time should be spent in discussion. Students must be consistently reminded that these are concepts, models, and explanations which explain observations. The only method of representing the structure of an atom is by use of higher mathematics which they are not in a position to use. They must not think of the atom as a tiny piece of matter which we can examine under a powerful microscope.

Tabulation of Results:



Sample graph for Experiment 6-2

See sample graph.

1. a. ${}_{2}$He
   ${}_{10}$Ne
   ${}_{18}$Ar
   ${}_{36}$Kr
   ${}_{54}$Xe

   b. ${}_{1}$H
   ${}_{9}$F
   ${}_{17}$Cl
   ${}_{35}$Br
   ${}_{53}$I (tied with ${}_{44}$Ru)

   c. ${}_{1}$H
   ${}_{3}$Li
   ${}_{11}$Na
   ${}_{19}$K
   ${}_{37}$Rb

2. See answer to Exp. 6-1, step 3 of Procedure.

Questions:

1. The periodicity of the pattern — the repetition within each section of the graph — should be emphasized. The recurring similarities in atomic structure are the basis for the periodic table which the students are building. This series of exercises is designed to point up this periodicity.

The periodicity occurs because of the pattern, or order, in which electrons enter the energy levels of atoms.

2. As atomic number increases within a vertical column, atomic size increases. This is because each element in the column has one more energy level than the preceding one. Addition of an energy level always results in an increase in the radius of the atom.

3. Within each section of the graph, as the atomic number increases, the atomic radius decreases in most cases. This is because electrons are being added to the same energy level, and the increased nuclear charge pulls them in closer.

4. As *d* orbitals fill, there is little change in atomic radius, because an inner energy level is being filled, and the radius is determined by the electrons in the outermost level.

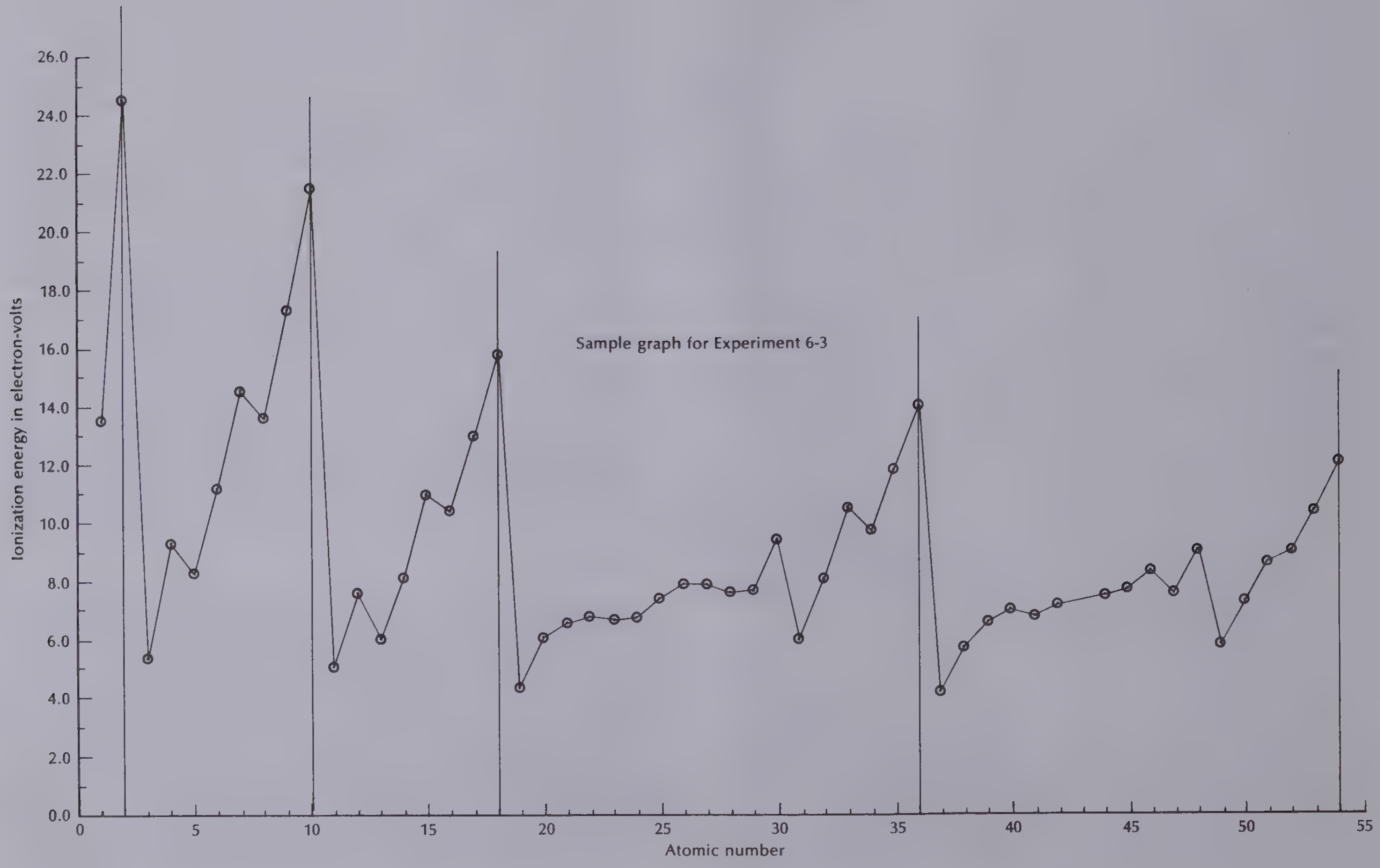
EXPERIMENT 6-3
THE RELATIONSHIP BETWEEN ATOMIC NUMBER AND IONIZATION ENERGY

Purposes:

To introduce another property of elements which is dependent upon the atomic number; to provide additional evidence that the properties of elements are recurring functions of the atomic number; to provide graphical evidence that the size of the atom influences the amount of energy required to remove the most loosely held electron.

Prelab Discussion:

The terms *ion* and *ionization energy* should be discussed so that students know their meanings prior to the graphing activity. When we speak of the ground state of an atom, we mean that the electrons are in their expected orbitals and the atom has minimum energy. When an atom absorbs energy and the electrons move to higher energy levels, we say the atom is excited. The excited state is unstable and the atom returns to its ground state by releasing energy.


Sample graph for Experiment 6-3

The ionization energy is the energy required to remove the least tightly held electron from an isolated atom in the gaseous state to form a positive ion. Ionization energy may be expressed in various units; the text gives values in electron volts. Do not go into these units now, except to say that they are simply units by which ionization energy is commonly measured. This activity requires only that students understand what is meant by an ion, and by ionization energy. Units of measurement can be considered in context at a later time.

Tabulation of Results:

See sample graph on previous page.

1. a.	b.	c.
$_{2}$He	$_{1}$H	$_{1}$H
$_{10}$Be	$_{3}$Li	$_{9}$F
$_{18}$Ar	$_{11}$Na	$_{17}$Cl
$_{36}$Kr	$_{19}$K	$_{35}$Br
$_{54}$Xe	$_{37}$Rb	$_{53}$I

Postlab Discussion:

Stress the periodicity, or regular recurrence, of similar patterns in the sections of the graph. Note especially that in the second and third sections, the pattern is almost identical; in the fourth and fifth the pattern is somewhat different from the first pattern, but almost identical for these two sections. In sections 4 and 5 we find *d* orbitals involved; there is very little difference in the ionization energies of the transition elements (those whose *d* orbitals are filling).

Placing this graph over the graph of atomic number versus atomic radius reveals an important fact: namely, that atoms with the largest radius have the lowest ionization energy. This is to be expected, because the farther an electron is from the nucleus, the less energy required to remove it. Remind students that the effective nuclear charge is closely related to the atomic number (number of protons in the nucleus). Conversely, the nearer the electron to the nucleus, the more energy required to remove it. These concepts will be of great importance when we begin our study of chemical bonding.

Conclusions:

1. Partly answered in the previous discussion. There is a sharp drop in ionization energy when an electron enters a new *s* orbital; there is a small drop in ionization energy when an electron enters the first *p* orbital (elements 5, 13, 31, 49); in three of the four cases there is a small drop for the fourth *p* electron (which has to pair up with one of the three electrons that are already in their own *p* orbitals). Except for the fourth *p* electron, there are rises in ionization energy as the *s* and *p* orbitals are filled further; there is a slight rise as the *d* orbitals are filled.

2. Within a vertical column, as atomic number increases, ionization energy decreases. (Reason: The outer electrons are further from the nucleus, so they are not as tightly held.) Within a section of the graph (energy level), as atomic number increases, so does ionization energy. (Reason: The outer electrons are in the same energy level, but the nuclear charge is greater, so the outer electrons are held more tightly to the atom.)

3. The peaks on one graph correspond to the valleys on the other graph. This indicates that the larger the atom, the easier it is to remove the most loosely held electron; the smaller the atom, the more difficult it is to remove the most loosely held electron.

4. As atomic number increases, the properties of elements recur at more or less regular intervals. This is because of the order of placing increasing numbers of electrons in the energy levels and orbitals.

5. Properties of the elements change in systematic and periodic ways as the atomic number changes.

EXPERIMENT 6-4
A CONVENIENT METHOD FOR GROUPING ELEMENTS

Purpose:

To introduce the student to the periodic table on the basis of his own studies which relate the properties of the elements to the atomic number.

Prelab Discussion:

Be sure students understand that they are to copy the vertical and horizontal columns of elements as they have them listed in previous exercises. It is very important that they cut out the columns as WHOLE columns. This makes shuffling them about much easier than if they had to work with 54 separate squares.

Postlab Discussion:

Students come out with a periodic table very much like the one they will be using during the remainder of the course. Students may place the column of rare gases (elements with the maximum number of electrons in the outer energy level) at the extreme left side of the chart. This is acceptable, because the numbers still read consecutively from left to right. However, it may be suggested that since these elements logically *complete* an energy level, it may be more convenient to let them end a horizontal row.

This is an ideal time to reveal the periodic table. The convential usage of the terms *group* and *family* as synonymous terms should be introduced. Also, it should be pointed out that the horizontal rows on the chart correspond to the filling of an energy level and that the conventional term for these rows is *period*.

Questions:

1. a. The theoretical electron configuration of element 110 ends with $7s^2\ 7p^2$. This places it in the column headed by ${}_{6}C$, two positions below ${}_{50}Sn$.
   b. Some assistance may be needed here. Since there are 14 elements in a row where *f* orbitals are filling, it becomes more convenient to place such rows below the rest of the chart.
2. (1) Elements with similar properties occur in vertical columns (for the most part).
   (2) Elements whose d orbitals are filling are readily located.
   (3) Relative atomic size may be obtained if one knows the trends.
   (4) If the trends are known, comparative ionization energies can be predicted.

EXPERIMENT 6-5
SEPARATION OF GROUP I CATIONS

Purposes:

To introduce the technique for separating and identifying the cations in Group I; to sustain interest while studying theory.

Materials: (per student)

10 cc of each of the following:
- distilled water
- 3M ammonia water (200 cc conc. ammonia water/liter of solution)
- 0.2M $AgNO_3$ (34 g/liter)
- 0.05M $Hg_2(NO_3)_2$ (28.1 g in 100 cc 6M HNO_3; dilute to 1 liter. Do *not* warm to dissolve; keep some free Hg in bottle to prevent oxidation to Hg^{2+})
- 0.2M $Pb(NO_3)_2$ (66 g/liter)
- 3M HCl (257 cc conc. acid/liter of solution)

5 cc of each of the following:
- 0.1M K_2CrO_4 (19.0 g/liter)
- 3M HNO_3 (192 cc conc. acid/liter of solution)

Prelab Discussion:

Explain the meaning of the terms *cation* and *anion*. Encourage students to think about the chemical reactions they observe and to try to explain these observations. Suggest that they write word equations for each reaction observed.

Questions:

1. Group I cations are precipitated by the chloride anion, Cl^-.
2. The group precipitate is treated with hot water in order to dissolve the lead chloride. This separates the lead chloride from the other chlorides which remain on the filter paper.
3. This black color is caused by a chemical reaction which forms free Hg, which is black, and $HgNH_2Cl$, which is white.
4. Ammonia water dissolves AgCl and places it in solution in the test tube. When sufficient HNO_3 is added to completely neutralize the base, the AgCl re-precipitates.

EXPERIMENT 6-6
IDENTIFICATION OF GROUP II CATIONS

Purposes:

To introduce the technique for separating and identifying the cations in Group II; to sustain interest while studying theory.

Materials: (per student)

10 cc of each of the following 0.2M solutions:
- $SbCl_3$ (45 g/liter; solvent, 3M HCl) Note: Make 3M HCl by mixing 257 cc of conc. acid with water to make one liter of solution.
- $Bi(NO_3)_3 \cdot 5H_2O$ (97 g/liter)
- $HgCl_2$ (54 g/liter)
- $Cu(NO_3)_2 \cdot 3H_2O$ (49 g/liter)

10 cc thioacetamide, 10% (100 g/liter. Note: This solution does not keep well and should be mixed in small quantities.)

We strongly urge the use of thioacetamide instead of the generation of poisonous H_2S by adding HCl to FeS. However, if you must use the older method, urge students to do the operation in a fume hood!

Prelab Discussion:

Little prelab discussion is needed because this experiment follows the same general procedure as that for the separation of Group I cations.

Questions:

1. Antimony sulfide—orange
   Copper(II) sulfide—black
   Bismuth sulfide—brown
   Mercury(II) sulfide—yellow black
2. Encourage students to recall the procedure used in Experiment 6-5. They should then reason that HCl would need to be added first in order to precipitate any Group I cations from solution. The Group II cations are soluble in HCl and will be caught

in the filtrate when the mixture is filtered. Then the filtrate can be tested with thioacetamide or H_2S as in this experiment.

3. The sulfides of Group II cations will precipitate only if the solution is slightly acid.

CHAPTER 7

CHEMICAL BONDING

Goals:

1. To introduce the concept of chemical bonding.
2. To introduce electronegativity and to use the periodic table of electronegativities to predict the types of bonds that can be expected to form when atoms unite.
3. To relate ionic size to melting temperature.
4. To relate potential energy to the formation of a covalent bond between identical atoms.
5. To obtain laboratory evidence for the relationship of bond types to solubility, chemical activity, and physical properties.

Teaching Hints:

1. We have found that students who are familiar with the trends in properties exhibited by elements in periods and families on the periodic table have little difficulty with the concepts of chemical bonding. Normally students who bog down during the study of bonding do not understand the use of the periodic table of electronegativity (Table 7-1) sufficiently well to apply generalizations.

Be alert to any problems each student may encounter and be prepared to assist him with the proper remedial work relating to the use of Table 7-1.

2. Encourage as much creative work as possible, especially in the area of free labs.

3. The film "Chemical Families" (a Chem Study film) is very helpful in providing information about the properties of related elements. It is highly recommended for viewing at this time.

4. Provide supplementary exercises as needed to give practice in predicting bond types and in predicting whether the melting points of specified compounds will be high or low. Students should consider electronegativity differences and molecular shapes in making their decisions.

Examples:

(Ionic compounds—have high melting points)

KCl, $CaCl_2$, $LiBr$, $MgCl_2$

(Molecular compounds with polar covalent bonds between atoms *within* the molecule but with molecules that are themselves nonpolar—have low melting points)

CO_2, NO_2, CCl_4, $SiCl_4$

EXPERIMENT 7-1
THE RELATIONSHIP BETWEEN BOND TYPES AND THE SOLUTION PROCESS

Purposes:

To determine the effect of bond types of both solute and solvent on the process of solution; to provide an opportunity for students to interpret observations by utilizing theoretical principles; to provide a foundation of experience for understanding the theory of solution.

Materials: (per student)

18 test tubes, 13 x 100 mm
1 test tube support
18 plastic squares (cut from plastic wrap)
18 labels
1 pair forceps
40 cc distilled water
40 cc ethanol, C_2H_5OH (denatured)
40 cc carbon tetrachloride, CCl_4
6 cc oil (cooking oil, light machine oil, or mineral oil)
1 teaspoon granulated sugar
1 teaspoon sodium chloride, $NaCl$
3 small iodine crystals, I_2

Prelab Discussion:

Stress the importance of labelling the test tubes so that observations on the three sets of mixtures can be compared. Just a pinch of sugar and salt are sufficient for each test. Using too much may give misleading results, because a saturated solution may form. This causes students to think that none of the solid dissolved.

Caution students not to handle I_2 with bare fingers. They should use forceps. Also, they should be reminded not to inhale CCl_4 vapors and to wash their hands immediately should they spill any CCl_4 on them.

Conclusions:

1. *Ionic*	*Polar covalent*	*Nonpolar covalent*
NaCl	water	CCl_4
	ethanol (?)	oil
	sugar (?)	ethanol (?)
		I_2
		sugar (?)

2. Solvents tend to dissolve solutes which have a bond type similar to their own. Hence the old expression, "Like dissolves like." If the bond types are very different, solution does not occur to any appreciable extent.

3. These observations may be valuable in assisting the student to explain the results of other experiments and to predict what may occur when a given solute is placed in a given solvent.

Postlab Discussion:

Results are fairly straightforward. Students may be puzzled about how to classify ethanol and sugar. They will observe that water, a very polar solvent, dissolves both ethanol and sugar quite well; and that I_2, which they know to be nonpolar, dissolves very well in ethanol. This should lead them to conclude that perhaps sugar and ethanol molecules have both polar and nonpolar ends. Students have not yet been introduced to the structure of these molecules and hence, this makes for a nice discovery experience. Encourage them to wonder and to propose possible hypotheses while they do further reading.

EXPERIMENT 7-2
THE RELATIONSHIP BETWEEN BOND TYPES AND PHYSICAL PROPERTIES

Purposes:

To provide another opportunity for students to study the relationship between bond types and properties of solids; to provide experience in making predictions based upon observations and knowledge of the periodic table of electronegativities (reprinted from text on p. 136 of lab manual).

Materials: (per student)

5 g NaCl
5 g paradichlorobenzene flakes
10 cc benzene
4 100 x 13 mm test tubes

Prelab Discussion:

Especially caution students to use a *very* small amount of paradichlorobenzene and *not* to inhale the vapors of paradichlorobenzene during the melting process. This test tube can be cleaned by rinsing it with a little benzene. Have students place the waste in a designated container to be disposed of later. It should not be poured into the sink. (If poured on the ground near ornamental shrubs it tends to deter damage by neighborhood dogs.)

Results:

Property	*Sodium chloride (Ionic)*	*Paradichlorobenzene (Molecular)*
Volatility (high or low)	low	high
Hardness	hard, granular	soft, waxy
Melting point (high or low)	high	low
Solubility in water	dissolves	does not dissolve
Solubility in benzene	does not dissolve	dissolves

Questions:

1. The stronger the forces holding the particles together in the solid, the less volatile the solid. If the intermolecular forces are weak, then the solid tends to be more volatile.
2. Water, dry ice.
3. When the bonds between particles are strong, as in ionic solids, the melting point is high because it requires a great deal of energy to overcome the attractive forces between the ions. When the forces holding particles together are weak, as in molecular solids, the melting point is low (relatively speaking). This is because less energy is required to overcome the weak van der Waals forces holding the molecules together.
4. Since a solution of sodium chloride is a good conductor of current, this indicates that ions are present. It would be expected to conduct a current if melted. On the basis of observations in this experiment, paradichlorobenzene in the liquid state would not be expected to conduct a current. This is because it is a molecular solid and no charged particles (ions) are present either in the solid or liquid state.
5. Sodium chloride, because it has the highest melting point. This indicates a great attraction between the ions of which it is composed.
6. The bonds holding ionic solids together appear to be much stronger than those holding molecular solids together.
7. a. sulfur : molecular
   b. potassium iodide : ionic
   c. calcium chloride : ionic
   d. iodine : molecular
   e. sulfur dioxide : molecular

EXPERIMENT 7-3
ENERGY IN CHEMICAL REACTIONS

Purposes:

To obtain experimental evidence that heat is evolved in a neutralization reaction; to provide an opportunity for the "discovery" of the law that heats of reaction add; to stress the mechanism in the solution process.

Materials: (per student)

2 Styrofoam cups (approximately 250 cc capacity)
1 thermometer
2 graduates (or volumetric flasks), 100 cc
200 cc distilled water
2 g solid NaOH
50 cc 0.5M HCl (42.8 cc conc. acid/liter)
50 cc 0.5M NaOH (20 g/liter)
100 cc 0.25M HCl (21.4 cc conc. acid/liter)

Prelab Discussion:

Remind students to handle solid NaOH with tweezers, *not* fingers! Also caution them that NaOH quickly absorbs moisture from the air and that if any is spilled to retrieve it immediately with tweezers and to wash the area with a piece of paper towel.

Solutions should be mixed as quickly as possible so that the temperature changes can be read correctly. Thermometers should be read at eye level.

A discussion of the calorie as a unit of measuring heat is usually necessary. Directions for calculating calories are given in the introduction, but may require some discussion. It is recommended that the better students read supplementary material on the Law of Additivity of heat reactions. Do not pursue this with slower students.

Questions:

1. Individual results vary, but if an average for the class is taken, it turns out that the amount of heat produced by the reaction of Part III is just about the same as the *sum* of the amounts of heat produced by the reactions in Parts I and II. This is because the reaction in Part I represents the heat of solution of NaOH; the reaction in Part II represents the heat of formation of H_2O from H_3O^+ + and OH^- ions; the reaction in Part III represents the combined heat of solution for solid NaOH plus the heat of formation of water from the hydronium and hydroxide ions.
2. A mole of NaOH weighs 40 g, so 1 gram = 0.025 mole.
3. The number of calories evolved per mole in each reaction will be the number of calories evolved for the experimental quantity (1 g) of NaOH divided by 0.025 mole. (Dividing by 0.025 is the same as multiplying by 40.)
4. Answers vary. However, if students take note of the titles for each part of the experiment, they should readily conclude that these are indicative of the source of heat.

CHAPTER 8

QUANTITATIVE RELATIONSHIPS IN CHEMISTRY

Goals:

1. To review the meaning of atomic weight.
2. To emphasize the usefulness of the mole concept.
3. To provide experience in formula writing.
4. To provide practice in writing chemical equations.
5. To use chemical equations for calculating quantitative relationships.
6. To provide experience in naming compounds.
7. To apply quantitative relationships in chemistry to laboratory work.

Teaching Hints:

1. This chapter may be a problem for students who are fearful of mathematics. The degree to which you require mastery of this material will necessarily depend upon the individual student. We recommend that those who are able to do so be encouraged to attain a high degree of proficiency with quantitative concepts, in writing chemical formulas and equations, and in naming compounds.

The use of charts containing the formulas for complex ions and the oxidation numbers of the elements and ions is recommended. Be sure students realize that Tables 8-1 and 8-4 in the chapter are supplemented by Appendix 6 (p. 295). Students need not memorize the information in these tables. The important point is that they learn how to use these aids to make their laboratory work more meaningful.

Appendix 5 (Formula Writing), Appendix 7 (Equation Writing), and Appendix 10A (Weight-Weight Relations) are also relevant to this chapter.

2. It is helpful to prepare practice exercises in quantity on varying levels of difficulty. These can then be given to individuals as they need them for drill. These should include practice in writing formulas, naming compounds, writing and balancing chemical equations, and in solving problems using the mole concept. These may be duplicated ahead of time in order that they might be available as needed.

3. The experiments that accompany this chapter have proved to be helpful in conveying the concepts that are presented in the textbook material. Specifically, they demonstrate evidence about the factors that affect the rate of a reaction (Experiments 8-1 and 8-2); require close attention to details (Experiment 8-3); and provide opportunities to observe and apply the results of those observations (Experiment 8-4).

EXPERIMENT 8-1
PREDICTING REACTION RATES

Purposes:

To provide experimental evidence that the concentration of a base affects the rate at which it reacts with aluminum metal; to provide experience in making and testing predictions based upon observations; to provide additional experience in graphing data.

Materials: (per student)

5 cc of each of the following solutions:
- 3M NaOH (120 g/liter)
- 6M NaOH (240 g/liter)
- 9M NaOH (360 g/liter)

1 2-cm x 3-cm piece of aluminum foil
steel wool
5 100 x 13 mm test tubes
stop watch or other timing device with a sweep second hand.

Prelab Discussion:

Remind students that NaOH solution is corrosive and should be handled with care. Caution them to rinse the outside of the reagent bottle after pouring so that the next person who picks it up will not get the solution on his hands. Tell students to record time from the instant the foil is dropped into the solution until it just disappears completely. Call attention to the fact that they are to make predictions for Part II based on observations in Part I.

Since students used aluminum foil in Experiment 2-2 where they rubbed it with steel wool, they may wish to introduce this variation into this experiment. Generally, the rubbed foil reacts more rapidly in a given concentration of NaOH solution. This is because there is usually a thin coating of aluminum oxide on the surface that requires more time to penetrate. However, this is not always evident.

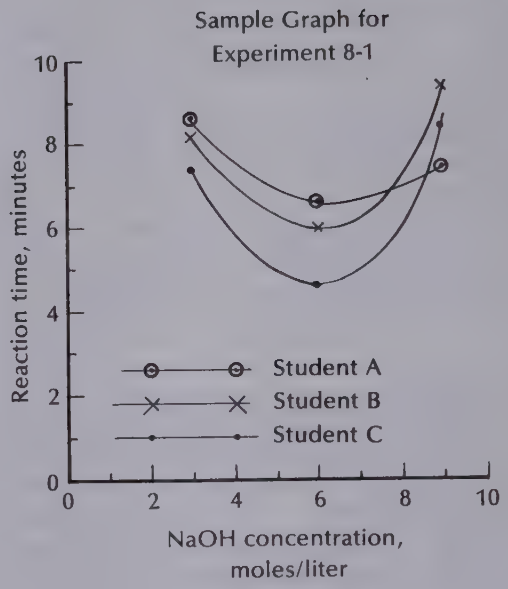
Results:

Students will probably predict that, compared to the reaction time for 6M NaOH, that for 3M NaOH will be longer and that for 9M NaOH will be shorter. However, the data will show otherwise. Sample data from three students were:

	Reaction time (min)		
	Student A	Student B	Student C
3M NaOH	8.6	8.2	7.4
6M NaOH	6.6	6.0	4.6
9M NaOH	7.4	9.4	8.4

Analysis of Data:

1. See sample graph.



2. Reaction times vary from student to student. Part of this variation is due to the fact that they may use different quantities of solution and foil strips of slightly different size. However, most students obtain results that exhibit similar *trends*. The trend is the important point to observe and analyze: The reaction time is shortest for the 6M concentration; weaker or stronger concentrations lengthen the reaction time.

3. This question poses some difficulties for students. Remind them that concentration is measured in moles per liter of solution. In a given liter of solution there are just so many water molecules. In the 3M NaOH solution there are apparently too few ions to attack the aluminum to produce a rapid change. With NaOH the 6M concentration seems to have the best balance between ions and water molecules, as evidenced by the most rapid reaction rate.

As the concentration of ions increases beyond 6M, the reaction becomes slower. Based upon our definition of a molar solution, we always have approximately the same amount of water in a liter of solution. Since more ions of the base are added as we increase the concentration to 9M, we might conclude that they are not being kept apart by the water molecules. In other words, as the number of Na^+ and OH^- ions increase, there are not enough water molecules to keep them from rejoining into NaOH ion-pairs. The presence of these ion pairs or even larger clusters may interfere with the opportunity for contact between the metal surface and the OH^- ions.

Since the differences in solution concentration are rather large, you could ask whether 6M NaOH is really the optimum concentration. It might be 5M or 7M. Encourage investigation of this possibility. Also encourage students to consider how many water molecules exist in the solution for each Na^+ and OH^- ion and how those water molecules would arrange themselves around these ions, especially at the various concentrations.

EXPERIMENT 8-2
THE EFFECT OF CONCENTRATION AND TEMPERATURE ON REACTION RATE

Purposes:

To provide quantitative evidence that the rate of a reaction depends upon the concentration of reactants and upon the temperature at which the reaction is conducted.

Materials: (per pair of students)

1 stopwatch (or timing device with sweep second hand)
2 50-cc beakers
1 250-cc beaker
2 10-cc graduated cylinders
12 18 x 150 mm test tubes
2 thermometers
150 cc solution B
110 cc solution A
50 cc distilled water
ice

Preparation of Solutions: (quantities for a class of 15 pairs)

Solution A (potassium iodate):
Preparing 2 liters of this solution will allow a little extra in case students need to repeat some trials. For 2 liters of solution, dissolve 14.33 grams of KIO_3 in sufficient water to make two liters.

Solution B (sodium hydrogen sulfite):
Preparing 2.5 liters of this solution will allow a little extra. 2.5 liters of solution requires:

0.5 g sodium metabisulfite, $Na_2S_2O_5$
10.0 g soluble starch
12.5 cc 1M H_2SO_4 (55.5 cc conc. acid/liter)

Make a paste of the starch by mixing it with a small amount of water. Add this paste slowly to 2400 cc boiling water, stirring all the while. Boil mixture for 5 minutes. Cool. Add other reagents to make 2.5 liters.

NOTE: If solutions are made up more than 24 hours prior to use, dissolve only the first two substances (sodium metabisulfate and starch) for solution B. Add the sulfuric acid immediately before use of the solution. Otherwise, the concentration of the HSO_3^- ion becomes too low to give good results.

Testing Solutions:

The solutions should be checked just prior to the experiment so that they can be adjusted to give appropriate reaction rates.

For Part I adjust the solutions so that between 10 and 15 seconds are required for reaction when the solutions are used in a 1:1 ratio at room temperature.

a. *Too much time:* If the reaction time is much longer than 15 seconds, add some $Na_2S_2O_5$ or a little more acid to solution B.

b. *Too little time:* If the reaction time is shorter than 10 seconds, dilute solution A.

For Part II dilute solution A to about one-half its original concentration in order to give a reaction rate at room temperature of about 25 seconds. If this is not done, the reaction rates at higher temperatures will be too rapid.

Recommendations:

If it can be managed, it is recommended that an entire class perform this experiment at the same time. This has several advantages: solutions do not keep well, and much time is conserved if they can be prepared for an entire group at one time. Also, instead of each pair of students having to do all dilutions, you can require that ALL do the 1:1 dilution and then assign three other dilutions to each pair. Class results can be pooled so that each student has a record of each dilution tested two or three different times.

The same is true of Part II. Let each pair do the room temperature reaction. Then assign two other temperatures and let them pool results.

Temperatures lower than about 5°C and higher than about 40°C give misleading results because of the instability of the solutions at these temperatures.

Prelab Discussion:

Assign dilutions and temperatures for which each pair of students is to be responsible. Indicate where the class results are to be reported so that all data will be available to each student.

Remind students that rapid pouring back and forth of the two solutions is essential to good results. Pouring two or three times is sufficient. One student should pour while the other records time.

Results:

Part I:

Solution B, cc	*Solution A, cc*		*Typical Reaction Time, sec*
10	10		13
10	9		15
10	8		18
10	7		20
10	6		24
10	5		30
10	4		43
10	3	It is not worthwhile doing these dilutions unless students want to try them.	55
10	2		59
10	1		65

Part II:

Temperature, °C	*Typical Reaction Time, sec*
5°	34
10°	30
15°	26
20°	23
25°	18
30°	16

Questions:

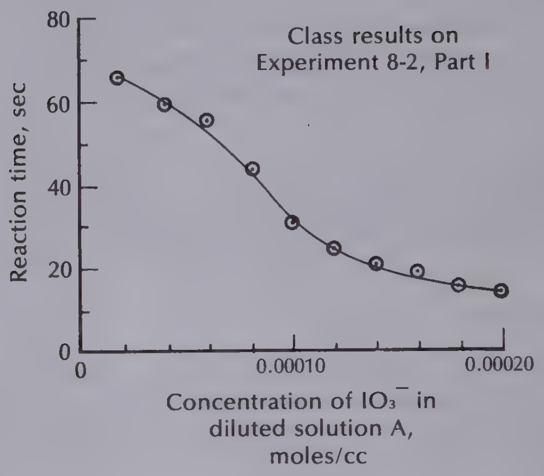
1. If more than one factor is varied at a time, it is difficult to explain the results. When the volume of solution A is kept constant, the effect on the reaction time of varying the concentration is more easily explained.

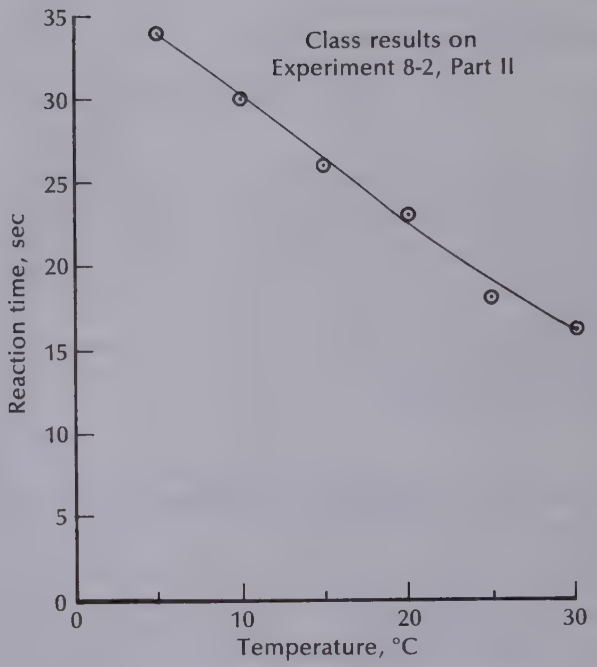
2. This may require some assistance, especially for weaker students. The initial concentration of KIO_3 is 0.02 M. This means that there are 0.02 moles of KIO_3 in one liter of solution. Since there are 1000 cc in one liter, then each cc of solution contains 0.00002 moles of KIO_3.

3.

Solution A	*Number of moles of KIO_3*
10 cc	0.00020
9 cc	0.00018
8 cc	0.00016
7 cc	0.00014
6 cc	0.00012
5 cc	0.00010
4 cc	0.00008
3 cc	0.00006
2 cc	0.00004
1 cc	0.00002

4,5. See sample graphs.





6. The more concentrated the solution, the shorter the reaction time. This means that the rate of the reaction is greater.

7. The higher the temperature, the shorter the reaction time and the greater the reaction rate.

EXPERIMENT 8-3
TYPES OF CHEMICAL REACTIONS

Purposes:

To provide opportunity for the student to observe examples of each of the four main types of chemical reactions; to provide experience in writing and balancing chemical equations; to provide opportunity for those who wish to gain experience in calculating quantitative relationships in chemical reactions; to strengthen student comprehension of the mole concept.

Materials: (per class, for demonstration in Part III)

For the demonstration on electrolysis of water, it is recommended that the Hoffman apparatus be used. This is large enough so that students can read the gas volumes and it can be used for a considerable length of time without needing to be set up again. If a power supply of 6 to 12 volts is used, it is easy to demonstrate the effect on the reaction rate of varying the voltage. This demonstration can be done in connection with the experiment, or if you prefer, it may be observed at a later time following the experiment.

(per student):

4 cc 0.2M $AgNO_3$ (34 g/liter)
4 cc 0.2M $CuSO_4 \cdot 5H_2O$ (50 g/liter)
copper foil, 0.5 cm x 3 cm
  (Copper wire can be used, but it must be unlacquered. A 6-cm length of #16 gauge works very well.)
mossy zinc, small piece
magnesium ribbon, 4 cm
HgO, a very tiny pinch
2 cc 0.5M $Pb(NO_3)_2$ (165 g/liter)
2 cc 0.5M KI (83 g/liter) (NOTE: NaI can be used equally well. For 0.5M NaI use 75 g/liter)
5 test tubes, 13 x 100 mm
1 asbestos mat
1 pair crucible tongs
1 wooden splint

Prelab Discussion:

Caution students not to get silver nitrate on their hands, as it stains. The importance of observing the reactions in Part I at intervals should be stressed. It is not necessary to watch them constantly, but they should be observed over a long enough period of time so that the change in color of the solutions is apparent to the student.

The test for oxygen gas should be discussed. In Part III students tend to insert the glowing splint before enough HgO has decomposed to give a test for oxygen. Help them to reach the appropriate conclusion: at first no oxygen is produced (or there is not enough to increase the glow); then, as the HgO is heated more strongly, or for several minutes longer, the test for oxygen is positive, and the appropriate interpretation can be made. Because of the possible presence of mercury vapor, should the upper portion of the test tube, near the mouth, be heated, this investigation is better carried out in a hood. In any case, caution students to avoid heating the upper portion of the test tube; caution them to avoid placing their noses near the mouth of the test tube. Even at room temperature, mercury evaporates and the vapors are severely harmful. Although the quantities involved here are not seriously hazardous, students' work should be closely supervised. The mercury that is produced in this experiment should be collected and kept in a closed container, such as a stoppered test tube or flask. If the container is well stoppered, there is no harm in allowing the students to exercise their natural curiosity about the unique properties of metallic mercury. If the container breaks and mercury is spilled, dust the area with copious quantities of powdered sulfur and then sweep up the mess (the mercury is rather quickly converted to HgS, which is harmless). Under no circumstances allow students to touch mercury with the skin, nor to amalgamate it with a penny (to make a "silver" penny) nor with a gold ring or other piece of personal jewelry.

If a binocular microscope can be set up so that students can observe the silver crystals closely (Part I), it adds interest. This part of the experiment should be related to Expt. 1-1, Part II.

Postlab Discussion:

Students will probably need help in completing the equations since this is their first exposure to such an activity. Stress the fact that the general rules for predicting products apply fairly well, but that these are generalizations and the only way to be sure of the products is through laboratory experimentation.

Many students will inquire whether the single replacement reactions will work in reverse. That is, will copper react with zinc sulfate? Encourage investigation of these questions in free labs rather than answering directly.

Much attention should be given to completing the word equations and to writing the formulas for reactants and products correctly by applying oxidation number rules and by proper balancing of equations by use of coefficients. Students should use the oxidation number tables (p. 137 of lab manual and Appendix 6 of text).

Equations:

1. copper + silver nitrate → copper (II) nitrate + silver
   $Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$
2. zinc + copper (II) sulfate → zinc sulfate + copper
   $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$
3. mercury (II) oxide → mercury + oxygen
   $2HgO \rightarrow 2Hg + O_2$
4. magnesium + oxygen → magnesium oxide
   $2Mg + O_2 \rightarrow 2MgO$
5. lead (II) nitrate + potassium iodide → lead (II) iodide + potassium nitrate
   $Pb(NO_3)_2 + 2KI \rightarrow 2KNO_3 + PbI_2$
6. water → hydrogen + oxygen
   $2H_2O \rightarrow 2H_2 + O_2$

7. aluminum + hydrochloric acid → aluminum chloride + hydrogen
$2Al + 6HCl \rightarrow 2AlCl_3 + 3H_2$
8. hydrogen + oxygen → water
$2H_2 + O_2 \rightarrow 2H_2O$
9. aluminum chloride + ammonium hydroxide → ammonium chloride + aluminum hydroxide
$AlCl_3 + 3NH_4OH \rightarrow 3NH_4Cl + Al(OH)_3$
10. iron + hydrochloric acid → iron (III) chloride + hydrogen
$2Fe + 6HCl \rightarrow 2FeCl_3 + 3H_2$
11. iron (II) sulfide + hydrochloric acid → iron (II) chloride + hydrogen sulfide
$2FeS + 4HCl \rightarrow 2FeCl_2 + 2H_2S$

Note: Some discussion of the different oxidation numbers of iron will probably be needed. For beginning students encourage them to use the most likely oxidation number by inspecting the equation. In equation 10 this is not critical, but in equation 11, iron has an oxidation number of +2 in iron sulfide, so they should use the same oxidation number in the product. Emphasize, however, that in many chemical reactions, elements do change oxidation number.

Questions:

1. Discuss reasons in terms of our definitions for the different types of reactions.

Reactions between metals and ions	*Synthesis reactions*	*Decomposition reactions*	*Rearrangement of ions*
1, 2, 7, 10	4, 8	3, 6	5, 9, 11

2. Part I: silver nitrate solution became blue-green; copper sulfate solution became colorless; silver crystals and copper metal were noted as the reaction progressed. Stress the fact that change in color of the solutions indicates that a different solution is being formed.

Part II: The most obvious evidence is production of a bright light. Heat is also evolved, but this probably will not be noticed. Product very different from magnesium in properties.

Part III: Gas is evolved; red powder changes to silver metal.

Part IV: Colorless solutions unite to form a yellow solid. Appearance of new material is evidence of chemical action.

3. Aluminum reacted with HCl to produce a gas; test tube became warm.

Hydrogen gas united with oxygen in the air with a popping sound.

Aluminum chloride reacted with ammonium hydroxide to produce a white, gelatinous solid. New substance indicates chemical reaction.

Iron reacted with HCl to produce a gas.

Iron sulfide reacted with HCl to produce a foul-smelling gas.

Optional:

These activities should be optional, but the more capable students will profit from the experience. Encourage those who will to do problems such as these at every opportunity. A sample problem is solved in Appendix 10 A (p. 301 of student text).

1. $Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$

Since 108 g of Ag is one mole, the number of moles in 10.8 g is 0.1 mole. We rewrite the equation:

$$??Cu + ??AgNO_3 \rightarrow ??Cu(NO_3)_2 + 0.1Ag$$

From the balanced equation we see that half as many moles of Cu are needed to produce a certain number of moles of Ag, so we need 0.5 X 0.1 mole, or 0.05 mole of Cu. Since one mole of Cu weighs 64 g, the amount of Cu needed is

$$0.05 \times 64 \text{ g} = 3.2 \text{ g}.$$

Since 10.8 g of Ag is 0.1 mole, and one mole contains 6.02 X 10^{23} atoms, there are 6.02 X 10^{22} atoms in 10.8 g of Ag.

(Note:
6.02 X 10^{22} = 60,200,000,000,000,000,000,000)

2. $2Mg + O_2 \rightarrow 2MgO$

There are the same number of moles of Mg and MgO, so 0.1 mole of magnesium oxide would be produced. One mole of MgO weighs 40 g, so 0.1 mole weighs 4.0 g. The number of molecules in 0.1 mole of MgO is the same as the number of atoms in 0.1 mole of Ag or the number of particles in 0.1 mole of anything: 6.02 X 10^{22}.

EXPERIMENT 8-4
A QUANTITATIVE STUDY OF A CHEMICAL REACTION

Purposes:

To provide laboratory experience requiring very careful attention to concentration of reactants, exacting techniques, and close observation; to enable students to derive the formula for a compound from the quantitative data obtained during its formation; to provide further experience in proposing hypotheses; to emphasize the mole concept.

Materials: (per student)

50 cc 0.5M KI (83.0 g/liter)
50 cc 0.5M $Pb(NO_3)_2$ (165.6 g/liter)
10 100 x 13 mm test tubes
2 10-cc graduated cylinders
1 250-cc beaker
1 test tube support
10 labels for test tubes
several plastic squares cut from plastic wrap
metric rules

NOTE: Since one major purpose of this experiment is for students to determine the ratio in which these reagents react completely, it is *most* important that the formulas do *not* appear on the reagent bottles. Use *names only*.

Prelab Discussion:

Warn students that unless they measure the volumes of the solutions very carefully, their results will not be sufficiently good to enable them to draw proper conclusions.

Cleanliness is of utmost importance in this experiment. Any contamination of one solution by the other results in a precipitate. Separate graduates must be used for each solution.

If plastic wrap (or a plastic garment bag) is cut into small squares, these make excellent covers for the test tubes while shaking to mix the solutions. This is better than using a stirring rod because less precipitate is lost.

Remind students to place labels on the test tubes as close to the lip as possible and to secure them tightly; otherwise, the labels will come off when the test tubes are placed in the beaker of water for heating.

The water in the 250 cc beaker should be adjusted to proper height *prior* to placing labels on the test tubes. If the student places 5 test tubes in the beaker and carefully adds water until it comes about ¾ the way up on the tubes, the labels should not get wet. The water should not be allowed to boil vigorously or else it will splatter over into the contents of the tubes.

Stress the importance of letting the precipitates settle evenly before measuring their heights. Measurement must be done carefully.

Urge students to make their predictions about the effect of adding a few more drops of the reagents to the test tubes *prior* to doing so. The important point here is this: They should add the extra reagent *one* drop at a time and observe very closely in order to determine if any more solid is forming.

Results are usually quite good when students carry out a careful performance of the experiment.

Sample Results:

Part I:

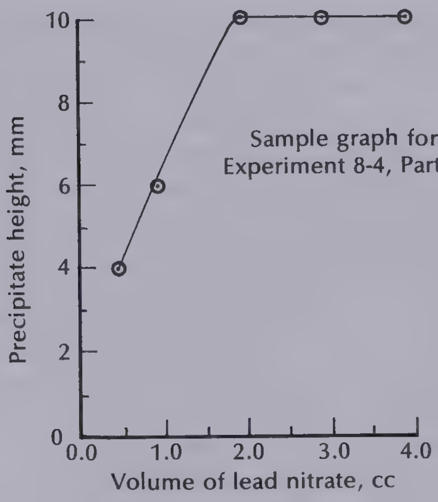
Tube	*KI, cc*	*$Pb(NO_3)_2$, cc*	*Precipitate height, mm*			
			Student A	B	C	D
1	4	0.5	4	3	5	4
2	4	1.0	6	7	6	6
3	4	2.0	10	9	10	9
4	4	3.0	10	11	10	10
5	4	4.0	10	10	11	9

Part II:

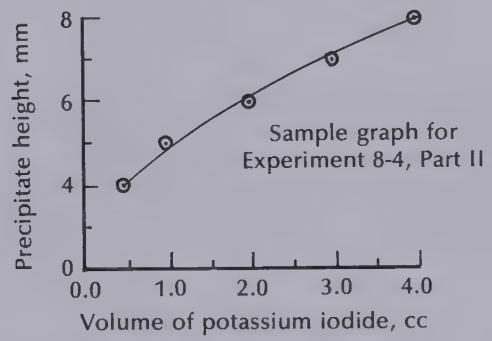
Tube	*$Pb(NO_3)_2$, cc*	*KI, cc*	*Precipitate height, mm*			
			Student A	B	C	D
1	2	0.5	4	3	5	4
2	2	1.0	5	4	6	6
3	2	2.0	6	5	8	7
4	2	3.0	7	6	9	8
5	2	4.0	8	7	9	9

Analysis of Data:

1, 2. See sample graphs.



Sample graph for Experiment 8-4, Part I

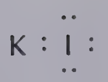


Sample graph for Experiment 8-4, Part II

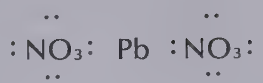
3. Within experimental limits, it appears that there must be twice the volume of potassium iodide as there is of lead nitrate — a 2:1 ratio.

4. potassium iodide + lead (II) nitrate →
potassium nitrate + lead (II) iodide.

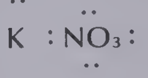
5. Note: This is the first time students encounter bonding between a two-element ion (NO_3^-) and another element. They need not be concerned with the bonds between the nitrogen and oxygen atoms but should realize that the nitrate ion has an electronegativity of its own (not stated in the text). On the basis of electronegativity, KI and PbI_2 should be polar covalent. However, both do conduct when dissolved in water, indicating they are actually ionic. Testing conductivity in water is a good way of determining the type of bond in the laboratory.

potassium iodide: polar covalent 

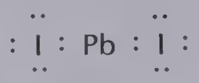
$\mathrm{K} : \overset{..}{\underset{..}{\mathrm{I}}} :$

lead (II) nitrate: ? 

$: \overset{..}{\underset{..}{\mathrm{NO_3}}} : \ \mathrm{Pb} \ : \overset{..}{\underset{..}{\mathrm{NO_3}}} :$

potassium nitrate: ? 

$\mathrm{K} \ : \overset{..}{\underset{..}{\mathrm{NO_3}}} :$

lead (II) iodide: polar covalent 

$: \overset{..}{\underset{..}{\mathrm{I}}} : \mathrm{Pb} : \overset{..}{\underset{..}{\mathrm{I}}} :$

6. Students can see from the electron dot diagrams that two iodide ions are required for each lead ion, whereas only one iodide ion is required for each potassium ion. This should help them to understand why two volumes of potassium iodide were required to react completely with one volume of lead nitrate. The point to emphasize here is this: the formula for a compound has an experimental basis. The mechanical writing of the formula by the use of oxidation numbers and the use of subscripts to make the net charge on the formula equal to zero is a convenient way of representing what we have learned from experimental evidence in the laboratory. This is the major point of the experiment.

7. If volumes were measured carefully, test tubes 1 and 2 probably contained unreacted iodide ions. Evidence to support this was obtained when a few more drops of lead nitrate were added to these tubes.

8. Test tubes 1, 2, 3, and 4. It was not until the ratio of potassium iodide to lead nitrate reached 2:1 that all lead ions were used up. It is very important to point out that this is our *theoretical* explanation. It is highly unlikely that we could ever measure so carefully that the exact ratio of ions themselves is 1:2. But within reasonable limits, our evidence indicates that our first statement is true.

9. The system was in equilibrium. Lead iodide was dissolving and crystallizing at equal rates, as indicated by the fact that no observable changes were taking place. The addition of more lead ions caused more lead iodide to form, because there were unreacted iodide ions in the solution.

A similar explanation applies to Part II.

10. If 4 cc of potassium iodide is required to react completely with 2 cc of lead nitrate, then it follows that any volume representing the 2:1 ratio would be required for a complete reaction. Therefore, two moles of potassium iodide are required to react completely with one mole of lead nitrate.

11. According to the electron dot diagram, for every mole of lead ions, two moles of iodide ions are needed. One mole of lead ions is supplied by one mole of lead nitrate, while two moles of iodide ions are supplied by two moles of potassium iodide.

12. $2KI + Pb(NO_3)_2 \rightarrow 2KNO_3 + PbI_2$

CHAPTER 9

PROPERTIES OF THE GASEOUS STATE

Goals:

1. To introduce the kinetic molecular theory to explain the properties of gases.
2. To apply the principles of bonding to explain the different condensation temperatures of substances.
3. To introduce standards for calculating the volume of gases.
4. To provide laboratory experiences with gases that will make the concepts meaningful to students.
5. To apply the effects of temperature and pressure on gases to daily activities.

Teaching Hints:

1. Experiment 9-1 is the first activity of the chapter. It should be performed prior to reading in the textbook.

2. Since Boyle's Law requires some experimental basis for understanding, it is recommended that the regular Boyle's Law J-tube apparatus be demonstrated *after* students have performed Experiment 9-2. This serves to reaffirm what they have discovered in the experiment and avoids the use of mercury by individual students. In the demonstration students can record and graph pressures and volumes.

3. Experiments 9-4 to 9-6 require extensive calculations. Some students will need considerable assistance in doing the calculations. Experiment 9-5 may be required of the more able students and made an optional activity for the others. It is an excellent exercise in making predictions based upon prior observations and is recommended for students who have sufficient backgrounds.

4. Continue to encourage creative work of all types, especially free labs.

5. If student error in Experiments 9-4 and 9-6 is unduly great, ask students to repeat the experiments. First, however, they should have their calculations checked for mathematical errors; often their work is satisfactory, but unusual mathematical errors creep into calculations. Again, the percentage of error permitted should certainly vary with the ability of each student. Hopefully, these experiments will assist students in understanding the concept of molar volume and the usefulness of its application in laboratory work.

EXPERIMENT 9-1
DIFFUSION OF GASES

Purposes:

To introduce the process of diffusion; to provide an example of a reaction between two gases.

Materials: (per pair of students)

2 ringstands
2 buret clamps
2 cotton plugs
1 glass or plastic tube, approximately 1 meter long and approximately 2 cm in diameter
1 meter stick
10 cc HCl, conc., in a stoppered container
10 cc ammonia water, conc., in a stoppered container

Prelab Discussion:

Call attention to the fact that gases move through the air. Remove the stopper from a bottle of perfume (or other more pungent material) and have students report when they note the odor. Raise the question as to what may determine the rate of motion of the gaseous molecules. Discuss setting up the apparatus and the importance of (a) having the tube level and (b) moistening the cotton plugs with the reagents simultaneously. Caution students about not getting HCl and ammonia water solutions on hands and clothing. Also they should keep the reagent bottles well apart.

Questions:

1. The white solid is ammonium chloride, NH_4Cl. The use of molecular models of HCl and NH_3 is quite useful here in leading the student to conclude what must happen when the two gases meet and react. This is an example of a synthesis reaction.

2. The first ring forms where the fastest moving HCl molecules meet the fastest moving NH_3 molecules and react to form NH_4Cl. The formation of the NH_4Cl creates a temporary barrier between the slower molecules of each gas and prevents a reaction between them when they arrive on the scene. However, the barrier breaks down when the NH_4Cl starts to settle out. This allows the two gases to meet and react, forming a second ring. The process is repeated over and over again.

3. On the average, molecules of NH_3 travel faster than molecules of HCl at the same temperature, because NH_3 has a lower molecular weight. Thus, the first ring forms closer to the HCl end of the tube. When the NH_4Cl barrier breaks down, there is a small empty region on either side of the first ring. Again, because the NH_3 molecules have higher average speeds than the HCl molecules, the two gases meet on the HCl side of the first ring. Each successive ring forms on the HCl side of the previous ring.

EXPERIMENT 9-2
THE EFFECT OF PRESSURE ON THE VOLUME OF GASES

Purpose:

To introduce students to Boyle's Law.

Materials: (per student)

10 to 25 cc plastic syringe
rubber stopper, #8
ringstand
2 buret clamps
5 or 6 identical books

It is recommended that transparent tape be placed over the graduations on the syringes because these rub off easily as students handle them.

Whether you bore out the rubber stopper or not is a matter of preference. We have obtained excellent results by using the solid stopper as a base for the syringe. In fact, unless the hole is bored so that the syringe fits very snugly into it, air escapes and poor results are obtained. If you do bore out a hole in the stopper, bore into the small end so that more stability is given to the apparatus.

Prelab Discussion:

Caution students to make the set-up as stable as possible. If the apparatus falls over, syringes may be broken and time lost. At least three trials should be made for each number of books used. The average value for each number of books is used for plotting the graph.

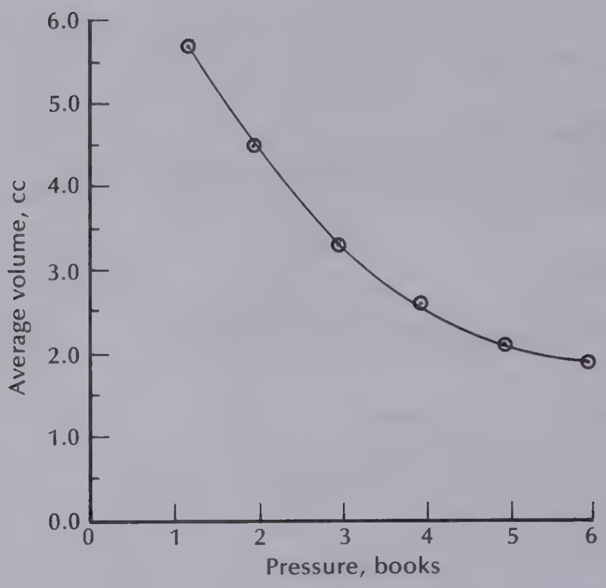
Results:

Sample student data:

No. of books	*Trial 1*	*Trial 2*	*Trial 3*	*Average volume in cc*
1	5.7	5.6	5.8	5.7
2	4.5	4.5	4.5	4.5
3	3.2	3.3	3.3	3.3
4	2.7	2.6	2.6	2.6
5	2.1	2.2	2.1	2.1
6	1.9	2.0	1.8	1.9

Analysis of Data:

1. See last column of sample data table.
2. See sample graph on following page.



3.

Average volume, cc	books	book · cc
5.7	1	5.7
4.5	2	9.0
3.3	3	9.9
2.6	4	10.4
2.1	5	10.5
1.9	6	11.4

4. At constant temperature, the volume of a gas decreases as the pressure increases.

5. As gases are compressed, the density increases. The same mass of material is crowded into less space, so the density (ratio of mass to volume) increases.

6. Reduce the pressure by one-half.

EXPERIMENT 9-3
THE EFFECT OF TEMPERATURE ON THE VOLUME OF A GAS

Purpose:

To introduce Charles' Law on the basis of experimental evidence.

Materials: (per student)

1 Charles' Law apparatus (see next column)
6 400-cc tall-form beakers
2 thermometers
1 ringstand
1 buret clamp
1 ring
1 asbestos mat
ice (cubes or crushed), about one beaker full

The Charles' Law apparatus can be prepared by sealing one end of a small bare glass tube and then introducing a mercury plug. To introduce the plug, warm the tube; while it is still hot, hold the open end in a small pool of mercury until about a one-cm length of mercury is in the tube. As the tube cools further, the plug will move into the tube. It is best for you or a lab assistant to prepare these tubes. Take care not to touch the mercury or breathe its vapors.

Prelab Discussion:

Best results are obtained when the trapped sample of gas is completely surrounded by the water bath.

The height of the column of gas is proportional to the gas volume (since the cross-sectional area of the tube is constant). Thus, students can record the height of the column, in cm, and consider these values to be values for the volume in "arbitrary units."

Caution the students against breathing any mercury vapors that might escape from the top of the tube. These vapors are poisonous.

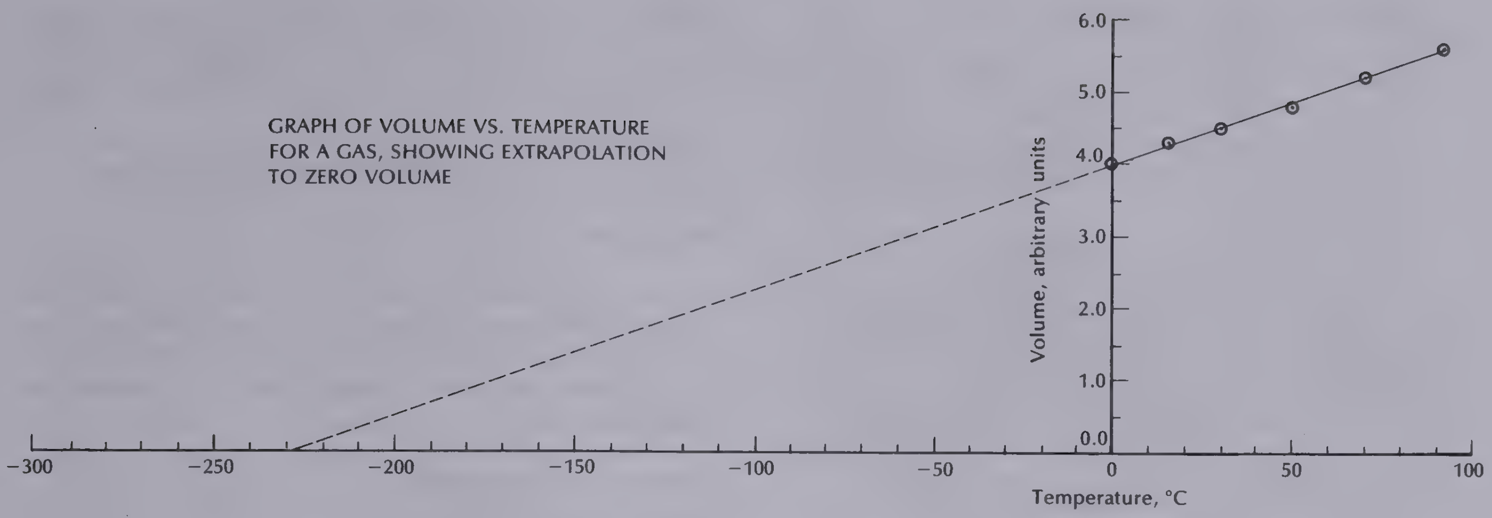
Demonstrate that heating a balloon filled with air causes the balloon to expand; cooling causes the balloon to become smaller. Raise a question about why this happens. Ask whether there may be a relationship between the volume of a gas and the number of degrees it is heated or cooled.

Postlab Discussion:

If a relationship between temperature and volume has not been found by the students, suggest they convert Celsius temperatures to Kelvin by adding 273°. The data give a straight-line graph when Celsius temperatures are used, but the direct proportion between *V* and *T* is not evident until the temperature conversion is made.

Sample Data:

Temperature, °C	*Volume, arbitrary units*
0	4.0
16	4.3
30	4.5
50	4.8
70	5.2
92	5.6


GRAPH OF VOLUME VS. TEMPERATURE FOR A GAS, SHOWING EXTRAPOLATION TO ZERO VOLUME

Analysis of Data:

1. See sample graph. The graph of volume *vs.* temperature is a straight line, showing that any change in temperature causes a change in volume that is proportional to the change in temperature.

2. According to the sample graph, the graph line crosses the temperature axis at −238°C. The theoretical value is −273°C, but student values will probably range between −200°C and −300°C. This temperature is the temperature at which the gas has zero volume. Theoretically, this is the lowest possible temperature, because the gas cannot have negative volume. Of course, real gases would become liquid at some temperature as they were cooled. But when graphs like this one are plotted for a variety of gases, over a wide range of temperatures, they all can be extrapolated to zero volume at −273°C (actually −273.15°C). If we take this value as the zero point on the Kelvin temperature scale, it becomes evident that the volume of a gas is directly proportional to its absolute temperature (the straight-line graph now goes through the origin for volume and temperature).

EXPERIMENT 9-4
THE MOLAR VOLUME OF A GAS

Purposes:

To introduce the concept of molar volume through experimental evidence; to provide practice in performing calculations.

Materials: (per student)

1 50-cc gas collecting tube
1 1-hole stopper to fit tube
1 large wide-mouth jar
10 cm fine copper wire to make cage
1 metric rule
1 thermometer
1 small piece of magnesium ribbon
1 ringstand
1 buret clamp
1 10-cc graduated cylinder
10 cc 6M HCl

The amount of magnesium ribbon to use is variable. It depends on the roll from which it is taken and on the atmospheric conditions of your locality. Determine ahead of time, by experiment, the length that will give between 40 and 50 cc of hydrogen under your circumstances.

Prelab Discussion:

Review definition for mole. Write both the word and the chemical equations for the reaction:

magnesium + hydrochloric acid →
magnesium chloride + hydrogen

$$Mg + 2HCl \rightarrow MgCl_2 + H_2$$

Remind students that one mole of magnesium produces one mole of hydrogen gas according to the equation. By knowing the weight of Mg used in the experiment, and by calculating the volume and pressure of the dry hydrogen gas produced, the volume of one mole of dry hydrogen gas at S. T. P. can be calculated.

Remind students to be careful not to let air into the gas collecting tube when inverting it. They should tap the tube gently before measuring the gas volume to insure that all the bubbles trapped below the meniscus of the solution are released.

Demonstrate how to level the water column inside the tube with the water in the jar. By placing their fingers over the holes in the stopper, students can transfer the gas tube to a jar to level the liquids. The jar must be deep enough to permit this.

Give students the weight of one meter of freshly opened magnesium ribbon.

Sample Data and Calculations:

1. Length of magnesium ribbon: 3.90 cm, or 0.039 m
2. Weight of magnesium ribbon:
   weight of one meter of Mg ribbon was 0.930 g
   weight of 0.039 m was therefore 0.039 X 0.930 g, or 0.0363 g
3. Number of moles of magnesium:
   weight of one mole of Mg = 24.3 g
   number of moles in 0.0363 g is therefore (0.0363 g) ÷ (24.3 g/mole) = 0.00149 mole
4. Barometric pressure: 675.8 torr
5. Partial pressure due to water vapor in tube: 36 torr at 32°C (interpolated from table in lab manual appendix, p. 132)
6. Partial pressure of hydrogen in tube:
   675.8 torr — 36 torr = 640 torr
7. Volume of hydrogen corrected to S.T.P.:
   volume of hydrogen at 32°C and 640 torr: 45.70 cc
   volume of hydrogen at S.T.P.

$$= 45.70\text{ cc} \times \frac{640\text{ torr}}{760\text{ torr}} \times \frac{273°\text{K}}{305°\text{K}}$$

$$= 34.4\text{ cc}$$

8. Volume of hydrogen produced by one mole of Mg:
   volume of H_2 produced by 0.00149 mole of Mg: 34.4 cc
   volume of H_2 produced by 1.00 mole of Mg is therefore:

$$34.4\text{ cc} \times \frac{1.00\text{ mole}}{0.00149\text{ mole}} = 23{,}000\text{ cc}$$

9. Experimental error:
   theoretical value: 22.4 liters = 22,400 cc
   experimental error = 23,000 cc − 22,400 cc = 600 cc
10. Percentage error:

$$\text{percentage error} = \frac{\text{experimental error}}{\text{theoretical value}} \times 100\%$$

$$= \frac{600\text{ cc}}{22{,}400\text{ cc}} \times 100\%$$

$$= 2\%$$

Questions:

1. At S.T.P. one mole of hydrogen gas weighs 2 grams and has a volume of 22,400 cc, so its density is

$$\frac{2\text{ grams}}{22{,}400\text{ cc}} = 0.00001\text{ g/cc}$$

2. The balanced equation is

$$Mg + 2HCl \rightarrow MgCl_2 + H_2$$

This can be interpreted as meaning that one mole of magnesium atoms produces one mole of hydrogen molecules (or two moles of hydrogen atoms).

EXPERIMENT 9-5
TESTING PREDICTIONS IN THE LABORATORY

Purpose:

To provide practice in making predictions about reactions which require students to use the results of previous experiments, their knowledge of laboratory techniques, and their imagination in planning an investigation.

Materials: (per student)

1 gas collecting tube
1 1-hole stopper to fit gas collecting tube
1 large wide-mouth jar
10 cm fine copper wire
1 10-cc graduated cylinder
1 ringstand
1 buret clamp
1 thermometer
60 cc 6M HCl
3 or 4 small pieces of zinc
3 or 4 small strips of magnesium ribbon
1 balance (readable to 0.01 g)

Gas collecting tubes larger than 50 or 100 cc can be made by using lengths of glass tubing with an interior diameter of one cm or larger. They can be plugged with a solid stopper and then calibrated by calculating the volume per linear cm or adding known amounts of liquid. The volumes can be marked along the tube on transparent tape.

Prelab Discussion:

Remind students that their plans for this experiment must be approved prior to doing any actual laboratory work. The major purpose of this experiment is to challenge them to use the results of Experiment 9-4 in planning the quantities of metals required in this activity. Their estimates should be checked, not so much for correctness, but to see if they have done good planning.

Sometimes the zinc does not react very readily with the HCl. If this happens, wrap the zinc metal with one or two loops of bare copper wire.

Results and Postlab Discussion:

The two chemical equations for the reactions in this experiment are:

$$Mg + 2HCl \rightarrow MgCl_2 + H_2$$
$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

The basic steps in making the prediction are as follows.

1. State volume of hydrogen desired, as measured at barometric pressure and room temperature.
2. Using table in lab manual appendix (p. 132), find partial pressure of hydrogen at this temperature.
3. Calculate the equivalent volume of the hydrogen at S.T.P.
4. Divide the volume in step 3 by 22,400 cc/mole to find the number of moles of hydrogen gas represented by that volume.
5. According to the balanced equations, the same number of moles of either Mg or Zn will produce a certain number of moles of H_2.
6. To find the weight of Mg or Zn required, multiply this number of moles by the weight of one mole (24.3 g for Mg, 65.4 g for Zn).

An important aspect of this experiment is that it provides the opportunity not only to make a prediction based on calculations but to test the prediction in the laboratory.

Questions:

1. The number of moles of hydrogen produced should be directly proportional to the weight of metal used, on each graph.
2. The number of moles of hydrogen produced equals the number of moles of metal used, according to the balanced equations. The number of moles of metal used is, of course, proportional to the weight of metal used.

EXPERIMENT 9-6
DETERMINATION OF THE FORMULA FOR AN UNKNOWN COMPOUND

Purpose:

To emphasize the practical application of the concepts of molar volume, gas laws, and percentage composition in determining the molecular formula for a compound.

Materials: (per student)

1 250-cc Erlenmeyer flask
1 1000-cc beaker
1 100-cc graduated cylinder
1 ringstand with buret clamp and ring
1 burner
1 5-cm square of aluminum foil
1 pin
balances
1 thermometer
5 cc "unknown" This can be ethanol, methanol, acetone, benzene or other liquid which boils well below 100°C. The percentage composition of the "unknown" must be given to students prior to the experiment.

	Percentage of each element		
Substance	*Carbon*	*Hydrogen*	*Oxygen*
ethanol	52.2	13.0	34.8
methanol	37.5	12.5	50.0
acetone	62.1	10.3	27.6
benzene	92.3	7.7	

Prelab Discussion:

Discuss what is meant by percentage composition. Go over the examples in Appendix 11 (pp. 303, 304 of the text). Show students that by knowing the percentage composition of a substance, one can derive the simplest formula (empirical formula) for the compound. Emphasize these steps:

(1) Divide the percentage of the element by its atomic weight.

(2) Divide each quotient by the smallest quotient to obtain a simple ratio of the number of atoms of each element in the compound.

(3) Write the simplest formula according to the results of Step 2.

The only way to know whether the simplest formula is the true (molecular) formula is to determine the molecular weight of the compound. Then, the empirical weight of the simplest formula is divided into the molecular weight; the molecular formula is obtained by multiplying each subscript in the simplest formula by the number of times the molecular weight is greater than the empirical weight of the simplest formula. (See example on p. 304.)

Caution students to be careful with boiling water. Remind them of the proper technique for pouring hot water.

Postlab Discussion:

Results on this experiment are usually quite good if the hole in the foil is very tiny and if the "unknown" is completely vaporized.

Sample Data and Calculations: (Calculations based on methanol)

1. Barometric pressure: 677.4 torr
2. Weight of flask, foil, and contents at end of experiment: 81.30 g
3. Weight of empty flask and foil: 81.00 g
4. Weight of unknown: 0.30 g
5. Volume of flask in cc: 267 cc
6. Volume of flask in liters: 0.267 l
7. Temperature of boiling water, observed: 96.0°C
8. Volume of vapor at S.T.P.: 0.176 l

(Original volume was 0.267 l at 96.0°C and 677.4 torr.)

9. Molecular weight of unknown compound:

$$0.30\text{ g} \times \frac{22.4\text{ l}}{0.176\text{ l}} = 38\text{ g}$$

10. Simplest formula of unknown compound: CH_4O

11. Molecular formula of unknown compound: CH_4O

38 is near enough to the empirical weight, 32, to indicate that the simplest formula and molecular formula are the same.

Questions:

1. The tiny hole insures that the flask will be just full of vapor at the time the unknown becomes completely gaseous. If the hole is large, too much of the unknown will escape. If there is no hole, the pressure inside the flask will be greater than barometric pressure.
2. Not allowing all the unknown to vaporize; permitting escape of too much unknown; having a mixture of air and unknown rather than a fairly pure unknown vapor.
3. This was the temperature of the gaseous unknown.
4. The vapor of the unknown is usually more dense than air. As the pressure in the flask increases, air molecules escape from the pinhole much more frequently than do molecules of the unknown. So, at the time when all the unknown has vaporized and the pressure in the flask equals barometric pressure, the flask is filled with unknown in vapor form.
5. The vapor pressure at that time is equal to barometric pressure.

Related Questions:

1. As the water in the beaker gets warmer, air bubbles form in the water. This air had been dissolved in the water, but comes out of solution because it is less soluble in hot water than in cold water.
2. Impurities or imperfections in the beaker cause bubble formation at these points. Have students test this by cleansing a beaker thoroughly with a detergent and rinsing well. Then pour a clear carbonated beverage into the beaker. Repeat, except this time, make a fingerprint at some place inside the beaker. Bubbles tend to form more easily where body oil has been deposited. Sprinkling salt or sugar into a carbonated beverage also provides seeds on which bubbles can build.

CHAPTER 10

PROPERTIES OF SOLIDS AND THEIR SOLUTIONS

Goals:

1. To introduce some of the properties associated with the solid state of matter.
2. To apply the principles of bonding to explain the properties of solids.
3. To apply Le Chatelier's Principle to changes in state.
4. To introduce a system for classifying crystalline solids on the basis of bond types.
5. To review the solution process, with special attention to the bonding in the water molecule.
6. To introduce ionic equations.
7. To provide laboratory experiences that will make these concepts more meaningful.

Teaching Hints:

1. Models of the different types of unit cells are very useful in helping students to understand the three common types.
2. Stress the importance of bond types in explaining the properties of solids.
3. Many of the concepts in this chapter have been introduced earlier in the course. The material in this chapter should serve to emphasize the usefulness of such ideas as bonding, equilibrium, and equation writing.
4. Encourage creative work. One form of such work is the construction of models of unit cells from Styrofoam spheres.

EXPERIMENT 10-1
WATER IN CRYSTALS

Purposes:
To study how crystalline solids differ in regard to water of hydration; to observe the properties of such solids before and after removing water of hydration; to propose possible explanations for the differences observed.

Materials: (per student)

small quantities of the following solids:
  sodium carbonate pentahydrate $Na_2CO_3{\cdot}5H_2O$
  sodium hydroxide NaOH
  sodium chloride NaCl
  calcium chloride, anhydrous $CaCl_2$
  potassium chloride KCl
  sodium chromate Na_2CrO_4
  copper (II) sulfate pentahydrate $CuSO_4{\cdot}5H_2O$
5 watch glasses
4 test tubes, 13 x 100 mm
1 test tube holder
1 balance

Prelab Discussion:

Caution students about how to handle NaOH. Also remind students that when they heat the $CuSO_4{\cdot}5H_2O$ crystal, the mouth of the test tube must be pointed slightly *downward;* otherwise, the water which condenses near the mouth may run back and cause the hot tube to break.

Results:

The following solids are deliquescent: NaOH, $CaCl_2$
The following solid is efflorescent: $Na_2CO_3 \cdot 5H_2O$

NaCl sometimes contains deliquescent impurities which may result in the noticeable absorption of moisture. However, pure NaCl is neither deliquescent nor efflorescent.

NaCl decrepitates; none of the other solids exhibit this property, though they may melt.

Blue $CuSO_4 \cdot 5H_2O$ crystals change to a white powder when the water of hydration is driven off completely (partial removal of the water yields a grayish-blue solid). The white powder immediately becomes blue when water is added.

Questions:

1. Keeping the lids on prevents the reagents from being ruined by loss of water of hydration or by the absorption of moisture from the air (depending upon the reagent present).
2. Deliquescent solids are used commercially to keep substances dry. Examples: in dessicators; in bottles containing $CoCl_2$ test paper. Machinery parts made of steel very often have a deliquescent substance packed with them to prevent rusting during shipment or storage.
3. The blue color is caused by the water of hydration.
4. The results with deliquescent solids will be much more apparent in very humid areas than in dry regions. The results with efflorescent solids will be more apparent in dry regions.
5. Common sodium chloride often contains a small amount of deliquescent magnesium chloride as an impurity.

EXPERIMENT 10-2
THE STRUCTURE OF CRYSTALS

Purposes:

To prepare and study ionic, molecular, and metallic crystals; to relate bond types to crystal structure.

Materials: (per student)

5 cc of each of the following saturated solutions:
  NaCl (21 g per 10 cc)
  $MgSO_4 \cdot 7H_2O$ (8 g per 10 cc)
  $NiSO_4 \cdot 7H_2O$ (8 g per 10 cc)
  KNO_3 (4 g per 10 cc)
5 cc 0.2M $AgNO_3$ (34 g/liter)
3 cc CS_2 (To be dispensed by the teacher or a responsible aide. This liquid is flammable; it is volatile and the vapors are flammable.)
3 cc 0.1M K_2CrO_4 (19 g/liter)
3 cc 0.1M $BaCl_2$ (24 g/liter)
small lump roll sulfur
½ teaspoon powdered sulfur
a few small iodine crystals
2 cm bare copper wire
5 glass slides
5 test tubes, 100 x 13 mm
1 50-cc beaker
1 evaporating dish
1 funnel and filter paper
1 magnifier (binocular microscope preferred)
1 ringstand with ring
1 asbestos mat
1 pair forceps

Prelab Discussion:

Caution students about the flammability of CS_2! Vapors should not be inhaled. Iodine should be handled with forceps, not bare fingers. Vapors should not be inhaled.

Crystals should be viewed through a good magnifier, preferably a binocular microscope.

Results:

Part I: sodium chloride: cubic
magnesium sulfate ($MgSO_4 \cdot 7H_2O$): rhombic or monoclinic
nickel sulfate ($NiSO_4 \cdot 7H_2O$): rhombic
potassium nitrate (KNO_3): rhombic
barium chromate ($BaCrO_4$): rhombic

Barium chromate is the precipitate that is formed when potassium chromate is mixed with barium chloride. It is generally in the form of such small crystals that they pass through filter paper. Heating causes the larger of the small crystals that are present to grow (the smaller crystals dissolve and reprecipitate on the larger ones). Then the solid can be removed by filtration.

Part II:

Roll sulfur dissolved in CS_2 forms rhombic crystals as the CS_2 evaporates. Melted sulfur forms monoclinic crystals upon cooling. Eventually they become rhombic, as this is the most stable form of sulfur at room temperature.

Rhombic iodine crystals form on the bottom of the evaporating dish as the vapors condense.

Part III:

Silver crystals form on the copper wire. These appear as needle-like structures. However, single silver crystals are cubic.

Questions:

1. Two ionic salts could be separated by evaporating the solution. The less soluble salt will crystallize first. This process is known as fractional crystallization.

2. Within a family on the periodic table, atomic radius increases as atomic number becomes larger. This is true because more energy levels must be occupied in order to accommodate the increasing numbers of electrons. As the atomic radius of the atoms becomes larger in molecular solids, the van der Waals forces become greater. This causes the molecules to become more difficult to separate. The melting point for each of the halogens is:

F_2: −220°C
Cl_2: −101°C
Br_2: −7.2°C
I_2: 114°C

3. Students may have already tried this as a free lab. If not, encourage them to do so. Copper crystals cannot be produced by placing a silver wire in a solution of copper sulfate. The reason is difficult to explain if students are not familiar with oxidation-reduction reactions. For the more able students, encourage them to read on this topic in reference books. For the less able, perhaps a sufficient explanation is that silver atoms are not able to take electrons from copper ions.

4. If solids pass through filter paper, this indicates that the particles are smaller than the pores in the paper.

5. Heating causes the larger crystals to grow at the "expense" of the smaller ones. The resulting particles of solid are too large to pass through the pores in the filter paper. (See earlier discussion of barium chromate under the Results section of this experiment.)

6. Ionically bonded crystals dissolve in water to make conductive solutions, although some are not very soluble. The bonds among the ions tend to be strong, leading to high melting points.

Molecular crystals have weak forces holding the molecules together in the crystal. This leads to low melting points. Unless the molecules are polar, they are not very soluble in water.

Metallic crystals (except for mercury) have high melting points. For additional discussion of bond properties, consult Chapter 10 of the text.

EXPERIMENT 10-3
IONIC REACTIONS IN WATER SOLUTION

Purposes:

To provide students with experience with precipitation reactions; to provide experimental evidence that certain combinations of ions form compounds which are not very soluble; to introduce the ionic equation; to provide sufficient experimental data so that students can formulate a set of solubility "rules" to serve as guidelines in future work; to provide practice in writing formulas and chemical equations.

Materials: (per student)

2 cc of each of the following solutions:
SET I: 0.1M $BaCl_2 \cdot 2H_2O$ (24 g/l)
0.1M Na_2SO_4, anhydrous (14 g/l)
0.1M $Ba(NO_3)_2$ (26 g/l)
0.2M $AgNO_3$ (34 g/l)
0.1M K_2CrO_4 (19 g/l)
0.1M $Pb(NO_3)_2$ (33 g/l)
SET II: $(NH_4)_2S$ (use liquid as it comes from supplier)
0.1M $CoCl_2 \cdot 6H_2O$ (24 g/l)
0.1M $Ni(NO_3)_2 \cdot 6H_2O$ (29 g/l)
0.1M $Cu(NO_3)_2 \cdot 3H_2O$ (24 g/l)
0.1M $FeCl_3 \cdot 6H_2O$ (27 g/l)
1M ammonia water (66.3 cc conc. solution/l)
SET III: 0.1M $Ba(NO_3)_2$ (26 g/l)
0.1M $CuCl_2 \cdot 2H_2O$ (17 g/l)
0.07M $AlCl_3 \cdot 6H_2O$ (17 g/l)
0.2M $AgNO_3$ (34 g/l)
0.1M $NiCl_2 \cdot 6H_2O$ (24 g/l)
0.1M Na_2CO_3 (10 g/l)

Note: Sets of solutions should be available for each student. Small plastic squeeze bottles with nozzles that permit the solution to come out in drops are recommended. Dropper pipets can be used, but there is a much greater chance of contaminating the solutions.

4 spot plates *or* sheet of clear plastic 15 cm X 15 cm, grease pencil, and pieces of black paper and white paper

Recommendations:

This is one of the most important experiments in the entire course because of the experience the students gain in observing, noting colors which certain ions impart to solutions, and in writing and balancing equations. It is extremely helpful to go over Set I with students after they have written the equations and *prior* to their moving on to another set. Encourage them to look at the products formed (these are all rearrangement reactions); to check the names of the compounds in the set solutions; to reason which product is likely to be the precipitate. For instance, if a precipitated product is thought to be the same as one of the solutions in a set, then it is instead obviously soluble in water. Hence, it can be eliminated and the other possible product must be the precipitate. In some cases, students will have soluble products with which they are not familiar and which are not members of one of the sets.

Encourage the students to withhold judgment on these until all tests have been made. Then they will be able to see that in no case has there been a precipitate with a K^+ ion, for instance.

This experiment is time consuming, but it is well worth the time if students reason their way through it — with help from you at appropriate times. Do not be ready with answers. Let them wonder; encourage them to hypothesize — this is one of the aims of the course!

Prelab Discussion:

Caution students about the danger of contaminating solutions. The tip of the dropper bottle should *never* touch the drop of solution already on the plate. Remind them that if a bottle is contaminated accidentally, to bring it to you so it can be replaced before other students obtain misleading results by using it.

Stress the fact that Set I will be completed and equations written prior to doing Sets II and III. Equations must be checked on Set I before students continue the experiment to insure that they are aware of what they are doing.

Some of the precipitates take a bit of time to become apparent; therefore, students should not rush their work. Colors of precipitates should be noted as these will be encountered again during the study of qualitative analysis.

Postlab Discussion:

SET I Eight different precipitates formed

	$Pb(NO_3)_2$	K_2CrO_4	$AgNO_3$	$Ba(NO_3)_2$	Na_2SO_4	$BaCl_2$
$BaCl_2$	white $PbCl_2$	yellow $BaCrO_4$	white $AgCl$	—	white $BaSO_4$	
Na_2SO_4	white $PbSO_4$	—	white Ag_2SO_4	white $BaSO_4$		
$Ba(NO_3)_2$	—	yellow $BaCrO_4$	—			
$AgNO_3$	—	rd-br. Ag_2CrO_4				
K_2CrO_4	yellow $PbCrO_4$					
$Pb(NO_3)_2$						

SET II Eight different precipitates formed

	NH_4OH	$FeCl_3$	$Cu(NO_3)_2$	$Ni(NO_3)_2$	$CoCl_2$	$(NH_4)_2S$
$(NH_4)_2S$	—	yl-gr. Fe_2S_3	black CuS	black NiS	black CoS	
$CoCl_2$	red $Co(OH)_2$	—	—	—		
$Ni(NO_3)_2$	green $Ni(OH)_2$	—	—			
$Cu(NO_3)_2$	blue $Cu(OH)_2$	—				
$FeCl_3$	rd-br. $Fe(OH)_3$					
NH_4OH						

SET III Six different precipitates formed

	Na_2CO_3	$NiCl_2$	$AgNO_3$	$AlCl_3$	$CuCl_2$	$Ba(NO_3)_2$
$Ba(NO_3)_2$	white $BaCO_3$	—	—	—	—	
$CuCl_2$	green $CuCO_3$	—	white $AgCl$	—		
$AlCl_3$	white $Al_2(CO_3)_3$	—	white $AgCl$			
$AgNO_3$	yellow Ag_2CO_3	white $AgCl$				
$NiCl_2$	green $NiCO_3$					
Na_2CO_3						

Summary of precipitate colors

White: $PbCl_2$, $PbSO_4$, $AgCl$, Ag_2SO_4, $BaSO_4$, $BaCO_3$, $Al_2(CO_3)_3$
Yellow: $PbCrO_4$, $BaCrO_4$, Ag_2CO_3
Red: $Co(OH)_2$
Red-brown: Ag_2CrO_4, $Fe(OH)_3$
Green: $Ni(OH)_2$, $CuCO_3$, $NiCO_3$
Blue: $Cu(OH)_2$
Black: CuS, NiS, CoS
Yellow-green: Fe_2S_3

Ionic equations for all reactions that produced a precipitate
Set I:

$$\begin{gathered} Ba^{2+}_{(aq)} + 2Cl^-_{(aq)} + Pb^{2+}_{(aq)} + 2NO_3{}^-_{(aq)} \rightarrow \\ Ba^{2+}_{(aq)} + 2NO_3{}^-_{(aq)} + PbCl_{2(s)} \text{ (white)} \end{gathered}$$

$$\begin{gathered} 2Na^+_{(aq)} + SO^{2-}_{(aq)} + Pb^{2+}_{(aq)} + 2NO^-_{3(aq)} \rightarrow \\ 2Na^+_{(aq)} + 2NO^-_{3(aq)} + PbSO_{4(s)} \text{ (white)} \end{gathered}$$

$$\begin{gathered} 2K^+_{(aq)} + CrO^{2-}_{4(aq)} + Pb^{2+}_{(aq)} + 2NO^-_{3(aq)} \rightarrow \\ 2K^+_{(aq)} + 2NO^-_{3(aq)} + PbCrO_{4(s)} \text{ (yellow)} \end{gathered}$$

$$\begin{gathered} Ba^{2+}_{(aq)} + 2Cl^-_{(aq)} + 2K^+_{(aq)} + CrO^{2-}_4 \rightarrow \\ 2K^+_{(aq)} + 2Cl^-_{(aq)} + BaCrO_{4(s)} \text{ (yellow)} \end{gathered}$$

$$\begin{gathered} Ba^{2+}_{(aq)} + 2NO^-_{3(aq)} + 2K^+_{(aq)} + CrO^{2-}_{4(aq)} \rightarrow \\ 2K^+_{(aq)} + 2NO^-_{3(aq)} + BaCrO_{4(s)} \text{ (yellow)} \end{gathered}$$

$$\begin{gathered} 2Ag^+_{(aq)} + 2NO^-_{3(aq)} + 2K^+_{(aq)} + CrO^{2-}_{4(aq)} \rightarrow \\ 2K^+_{(aq)} + 2NO^-_{3(aq)} + Ag_2CrO_{4(s)} \text{ (red-brown)} \end{gathered}$$

$$\begin{gathered} Ba^{2+}_{(aq)} + 2Cl^-_{(aq)} + 2Ag^+_{(aq)} + 2NO^-_{3(aq)} \rightarrow \\ Ba^{2+}_{(aq)} + 2NO^-_{3(aq)} + 2AgCl_{(s)} \text{ (white)} \end{gathered}$$

$$\begin{gathered} 2Na^+_{(aq)} + SO^{2-}_{4(aq)} + 2Ag^+_{(aq)} + 2NO^-_{3(aq)} \rightarrow \\ 2Na^+_{(aq)} + 2NO^-_{3(aq)} + Ag_2SO_{4(s)} \text{ (white)} \end{gathered}$$

$$\begin{gathered} 2Na^+_{(aq)} + SO^{2-}_{4(aq)} + Ba^{2+}_{(aq)} + 2NO^-_{3(aq)} \rightarrow \\ 2Na^+_{(aq)} + 2NO^-{}_{3(aq)} + BaSO_{4(s)} \text{ (white)} \end{gathered}$$

$$\begin{gathered} Ba^{2+}_{(aq)} + 2Cl^-_{(aq)} + 2Na^+_{(aq)} + SO^{2-}_{4(aq)} \rightarrow \\ 2Na^+_{(aq)} + Cl^-_{(aq)} + BaSO_{4(s)} \text{ (white)} \end{gathered}$$

Set II:

$$Co^{2+}_{(aq)} + 2Cl^{-}_{(aq)} + 2NH^{+}_{4(aq)} + 2OH^{-}_{(aq)} \rightarrow 2NH^{+}_{4(aq)} + 2Cl^{-}_{(aq)} + Co(OH)_{2(s)} \text{ (red)}$$

$$Ni^{2+}_{(aq)} + 2NO^{-}_{3(aq)} + 2NH^{+}_{4(aq)} + 2OH^{-}_{(aq)} \rightarrow 2NH^{+}_{4(aq)} + 2NO^{-}_{3(aq)} + Ni(OH)_{2(s)} \text{ (green)}$$

$$Cu^{2+}_{(aq)} + 2NO^{-}_{3(aq)} + 2NH^{+}_{4(aq)} + 2OH^{-}_{(aq)} \rightarrow 2NH^{+}_{4(aq)} + 2NO^{-}_{3(aq)} + Cu(OH)_{2(s)} \text{ (blue)}$$

$$Fe^{3+}_{(aq)} + 3Cl^{-}_{(aq)} + 3NH^{+}_{4(aq)} + 3OH^{-}_{(aq)} \rightarrow 3NH^{+}_{4(aq)} + 3Cl^{-}_{(aq)} + Fe(OH)_{3(s)} \text{ (red-brown)}$$

$$6NH^{+}_{4(aq)} + 3S^{2-}_{(aq)} + 2Fe^{3+}_{(aq)} + 6Cl^{-}_{(aq)} \rightarrow 6NH^{+}_{4(aq)} + 6Cl^{-}_{(aq)} + Fe_2S_{3(s)} \text{ (yellow-green)}$$

$$2NH^{+}_{4(aq)} + S^{2-}_{(aq)} + Cu^{2+}_{(aq)} + 2NO^{-}_{3(aq)} \rightarrow 2NH^{+}_{4(aq)} + 2NO^{-}_{3(aq)} + CuS_{(s)} \text{ (black)}$$

$$2NH^{+}_{4(aq)} + S^{2-}_{(aq)} + Ni^{2+}_{(aq)} + 2NO^{-}_{3(aq)} \rightarrow 2NH^{+}_{4(aq)} + 2NO^{-}_{3(aq)} + NiS_{(s)} \text{ (black)}$$

$$2NH^{+}_{4(aq)} + S^{2-}_{(aq)} + Co^{2+}_{(aq)} + 2Cl^{-}_{(aq)} \rightarrow 2NH^{+}_{4(aq)} + 2Cl^{-}_{(aq)} + CoS_{(s)} \text{ (black)}$$

Set III:

$$Ba^{2+}_{(aq)} + 2NO^{-}_{3(aq)} + 2Na^{+}_{(aq)} + CO^{2-}_{3(aq)} \rightarrow 2Na^{+}_{(aq)} + 2NO^{-}_{3(aq)} + BaCO_{3(s)} \text{ (white)}$$

$$Cu^{2+}_{(aq)} + 2Cl^{-}_{(aq)} + 2Na^{+}_{(aq)} + CO^{2-}_{3(aq)} \rightarrow 2Na^{+}_{(aq)} + 2Cl^{-}_{(aq)} + CuCO_{3(s)} \text{ (green)}$$

$$2Al^{3+}_{(aq)} + 6Cl^{-}_{(aq)} + 6Na^{+}_{(aq)} + 3CO^{2-}_{3(aq)} \rightarrow 6Na^{+}_{(aq)} + 6Cl^{-}_{(aq)} + Al_2(CO_3)_{3(s)} \text{ (white)}$$

$$2Ag^{+}_{(aq)} + 2NO^{+}_{3(aq)} + 2Na^{+}_{(aq)} + CO^{2-}_{3(aq)} \rightarrow 2Na^{+}_{(aq)} + 2NO^{+}_{3(aq)} + Ag_2CO_{3(s)} \text{ (yellow)}$$

$$Ni^{2+}_{(aq)} + 2Cl^{2-}_{(aq)} + 2Na^{+}_{(aq)} + CO^{2-}_{3(aq)} \rightarrow 2Na^{+}_{(aq)} + 2Cl^{-}_{(aq)} + NiCO_{3(s)} \text{ (green)}$$

$$2Ag^{+}_{(aq)} + 2NO^{-}_{3(aq)} + Ni^{2+}_{(aq)} + 2Cl^{-}_{(aq)} \rightarrow Ni^{2+}_{(aq)} + 2NO^{-}_{3(aq)} + 2AgCl_{(s)} \text{ (white)}$$

$$Cu^{2+}_{(aq)} + 2Cl^{-}_{(aq)} + 2Ag^{+}_{(aq)} + 2NO^{-}_{3(aq)} \rightarrow Cu^{2+}_{(aq)} + 2NO^{-}_{3(aq)} + 2AgCl_{(s)} \text{ (white)}$$

$$Al^{3+}_{(aq)} + 3Cl^{-}_{(aq)} + 3Ag^{+}_{(aq)} + 3NO^{-}_{3(aq)} \rightarrow Al^{3+}_{(aq)} + 3NO^{-}_{3(aq)} + AgCl_{(s)} \text{ (white)}$$

Analysis of Data:

1. Tabulation of ions forming precipitates.

Positive Ion	*Precipitates*
Ba^{2+}	$BaCrO_4$, $BaSO_4$, $BaCO_3$
Na^+	none
Ag^+	$AgCl$, Ag_2SO_3, Ag_2CrO_4, Ag_2CO_3
K^+	none
Pb^{2+}	$PbCl_2$, $PbSO_4$, $PbCrO_4$
NH_4^+	none
Co^{2+}	$Co(OH)_2$, CoS
Ni^{2+}	$Ni(OH)_2$, $NiCO_3$, NiS
Cu^{2+}	CuS, $Cu(OH)_2$, $CuCO_3$
Fe^{3+}	$Fe(OH)_3$, Fe_2S_3
Al^{3+}	$Al_2(CO_3)_3$

Negative Ion	*Precipitates*
Cl^-	$PbCl_2$, $AgCl$
SO_4^{2-}	Ag_2SO_4, $BaSO_4$, $PbSO_4$
NO_3^-	none
CrO_4^{2-}	$PbCrO_4$, Ag_2CrO_4, $BaCrO_4$
S^{2-}	Fe_2S_3, CuS, NiS, CoS
OH^-	$Co(OH)_2$, $Ni(OH)_2$, $Cu(OH)_2$, $Fe(OH)_3$
CO_3^{2-}	$BaCO_3$, $CuCO_3$, $Al_2(CO_3)_3$, Ag_2CO_3, $NiCO_3$

2. Potassium, sodium, and ammonium compounds are soluble.

Nitrates are soluble.

Chlorides are soluble except silver and lead chlorides.

Sulfates are soluble except lead, silver, and barium sulfates.

Carbonates are insoluble except sodium carbonate.

Hydroxides are insoluble except ammonium hydroxide.

Chromates are insoluble except sodium and potassium chromates.

NOTE: Students may state solubility rules in a wide variety of ways. The main point to this experiment is that they make proper generalizations on the basis of their *limited* evidence. Do *not* expect them to state the "rules" with which you are familiar. So long as they express the correct idea, they should be given credit.

3. There is no sure way to predict the color of compounds because both the anion and cation affect it. However, yellow solutions are likely to contain iron, chromate, or dichromate ions. Blue or blue-green solutions are likely to contain nickel or copper or chromium ions. Pink solutions are likely to contain cobalt or manganese ions.

CHAPTER 11

ELEMENTARY QUALITATIVE ANALYSIS

Goals:

1. To provide experience in separating and identifying cations and anions in water solutions.
2. To provide practice in interpreting results in the laboratory.
3. To provide additional practice in writing chemical equations to represent chemical reactions.
4. To apply the principles of solution, acid-base reactions, the common ion, and ionic theory to laboratory work.

Teaching Hints:

1. This chapter should be required of all students. Those who are behind might be told to "jump ahead" in order to work with an exciting part of chemistry. The more able students will find it extremely useful in tying together all that has been learned in the course in a meaningful and interesting way. The less able students enjoy the excitement of working on "unknowns." The extent that you require each student to write equations representing the reactions must be determined by the ability of the individual.
2. It is better to have all students work on this chapter at the same time and for a given block of time.
3. Much solution preparation is necessary for qualitative analysis.
4. There is a great need for the teacher to be readily available to the students to answer questions, receive and listen to reports, and to keep things moving along.
5. The importance of keeping good records should be stressed from the beginning.
6. Each student should complete Experiment 11-1 prior to beginning 11-2, as this provides experience with known tests. Review of Experiments 6-5 and 6-6 is recommended. Again, we strongly recommend the use of thioacetamide in place of hydrogen sulfide in testing for Group II cations. Hydrogen sulfide should not be used unless a fume hood is available and is in operating condition.

EXPERIMENT 11-1
IDENTIFICATION OF CERTAIN ANIONS

Purposes:

To introduce the techniques of separating and identifying common anions; to provide practice in writing chemical equations.

Materials: (per student)

about 10 cc of each of the following 0.1M solutions:
- $Na_2Co_3 \cdot 10H_2O$ (28.6 g/l)
- $NaNO_2$ (6.9 g/l)
- $Na_2SO_3 \cdot 7H_2O$ (25 g/l)
- Na_2S (7.8 g/l)
- K_2CrO_4 (19.4 g/l)
- $Na_3PO_4 \cdot 12H_2O$ (38 g/l)
- $Na_2SO_4 \cdot 10H_2O$ (32 g/l)
- $NaCl$ (6 g/l)
- $NaBr$ (10.3 g/l)
- NaI (15 g/l)
- $NaNO_3$ (8.5 g/l)
- $NaC_2H_3O_2$ (8.2 g/l)
- $NaClO_3$ (10.6 g/l)

about 5 cc of each of the following solutions:
- 3M HCl (256.5 cc conc. acid/liter)
- 0.2M $AgNO_3$ (34 g/l)
- 6M H_2SO_4 (336 cc conc. acid/liter)
- conc. H_2SO_4
- 0.5M $BaCl_2 \cdot 2H_2O$ (122 g/l)
- 6M HNO_3 (384 cc conc. acid/liter)
- 6M ammonia water (398 cc conc. base/liter)
- Cl_2 water (250 cc Clorox (or Purex) + 500 cc distilled water + 35 cc 6M HCl; dilute to 1 liter)
- CCl_4
- $FeSO_4$ (must be freshly prepared)

10 to 12 test tubes, 100 X 13 mm

Prelab Discussion:

Remind students of the meaning of *anion* and *cation*. Point out that this experiment is analogous to Experiments 6-5 and 6-6 (Identification of Group I and II Cations).

Questions:

1. The sulfite ion can be identified by the odor of sulfur dioxide. The acetate ion can be identified by the odor of vinegar.
2. The chlorate ion is identified by a green gas. The nitrite ion is identified by the reddish-brown color of nitrogen dioxide gas.
3. The bromide ion is identified by an orange CCl_4 layer. The iodide ion is identified by a violet CCl_4 layer.
4. The bromide or iodide ions lose electrons to the chlorine molecules and become bromine or iodine molecules, respectively. Solutions of bromine or iodine in CCl_4 have characterisitic colors, which are used for identification.
5. When HCl is added to the white precipitate, the precipitate will dissolve if it is phosphate, but will not dissolve if it is sulfate.
6. AgCl is a white precipitate, while AgBr and AgI are yellow precipitates.

Equations:

1. $2Na^{+}_{(aq)} + CO^{2-}_{3(aq)} + 2H_3O^{+}_{(aq)} + 2Cl^{-}_{(aq)} \rightarrow$
$2Na^{+}_{(aq)} + 2Cl^{-}_{(aq)} + 2H_2O_{(l)} + CO_{2(g)}$

2. $2Na^{+}_{(aq)} + 2C_2H_3O^{-}_{2} + 2H_3O^{+}_{(aq)} + SO^{2-}_{4(aq)} \rightarrow$
$2Na^{+}_{(aq)} + SO^{2-}_{4(aq)} + 2H_2O_{(l)} + 2HC_2H_3O_{2(l)}$

3. $3Na^{+}_{(aq)} + 3NO^{-}_{2(aq)} + 2H_3O^{+}_{(aq)} + SO^{2-}_{4(aq)} \rightarrow$
$3Na^{+}_{(aq)} + SO^{2-}_{4(aq)} + 2NO_{(g)} + NO^{-}_{3(aq)} + 3H_2O_{(l)}$
$2NO_{(g)} + O_{2(g)} \rightarrow 2NO_{2(g)}$

4. $2Na^{+}_{(aq)} + S^{2-}_{(aq)} + 2H_3O^{+}_{(aq)} + 2Cl^{-}_{(aq)} \rightarrow$
$2Na^{+}_{(aq)} + 2Cl^{-}_{(aq)} + 2H_2O_{(l)} + 2H_2S_{(g)}$

5. $2Na^{+}_{(aq)} + SO^{2-}_{3(aq)} + 2H_3O^{+}_{(aq)} + SO^{2-}_{4(aq)} \rightarrow$
$2Na^{+}_{(aq)} + SO^{2-}_{4(aq)} + 3H_2O_{(l)} + SO_{2(g)}$

6. $2Na^{+}_{(aq)} + SO^{2-}_{4(aq)} + Ba^{2+}_{(aq)} + 2Cl^{-}_{(aq)} \rightarrow$
$2Na^{+}_{(aq)} + 2Cl^{-}_{(aq)} + BaSO_{4(s)}$

7. $6Na^{+}_{(aq)} + 2PO^{3-}_{4(aq)} + 3Ba^{2+}_{(aq)} + 6Cl^{-}_{(aq)} \rightarrow$
$6Na^{+}_{(aq)} + 6Cl^{-}_{(aq)} + Ba_3(PO_4)_{2(s)}$

8. $2K^{+}_{(aq)} + CrO^{2-}_{4(aq)} + Ba^{2+}_{(aq)} + 2Cl^{-}_{(aq)} \rightarrow$
$2K^{+}_{(aq)} + 2Cl^{-}_{(aq)} + BaCrO_{4(s)}$

9. $Na^{+}_{(aq)} + Cl^{-}_{(aq)} + Ag^{+}_{(aq)} + NO^{-}_{3(aq)} \rightarrow$
$Na^{+}_{(aq)} + NO^{-}_{3(aq)} + AgCl_{(s)}$

EXPERIMENT 11-2
ELEMENTARY QUALITATIVE ANALYSIS

Purposes:

To close the study of beginning chemistry with laboratory experiences that will employ many of the concepts and principles learned throughout the course; to provide experiences in applying concepts to actual laboratory work; to provide experience in equation writing.

Materials: (per class of 24 students)

It is recommended that most of the liquid reagents be put in 250 cc bottles and arranged alphabetically on a table near which or at which students will work. One set of reagents per pair of students is desirable. Stock bottles of each reagent can be made up and maintained; the smaller bottles can be replenished from the stock bottles as needed.

Platinum wires, blowpipes, and charcoal blocks also may be placed on a central supply table where students have easy access to them.

Acids:
6M $HC_2H_3O_2$ (341.4 cc conc. acid/liter)
3M HCl (256.5 cc conc. acid/liter)
6M HNO_3 (384 cc conc. acid/liter)
1.5M H_2SO_4 (84 cc conc. acid/liter)
6M H_2SO_4 (336 cc conc. acid/liter)
Concentrated H_2SO_4

Bases:
6M ammonia water (398 cc conc. base/liter)
6M NaOH (240 g/liter)

Solids:
borax
iron(II) sulfate (to prepare fresh solution)

Liquid Regeants:
2M $(NH_4)_2CO_3$ (dissolve 192 g in 400 cc of 3M ammonia water and then dilute to 1 liter with water)
4M NH_4Cl (214 g/liter)
0.25M $(NH_4)_2C_2O_4 \cdot H_2O$ (35 g/liter)
$(NH_4)_2S$ (1 volume reagent grade ammonium sulfide, 1 volume conc. ammonia water, and 1 volume water)
0.5M $BaCl_2 \cdot 2H_2O$ (122 g/liter)
CCl_4
Cl_2 water (250 cc Clorox (or Purex) + 500 cc distilled water + 35 cc 6M HCl; dilute to 1 liter)
0.1M $Ca(NO_3)_2 \cdot 4H_2O$ (23.6 g/liter)
0.5M K_2CrO_4 (97 g/liter)
0.5M $K_4Fe(CN)_6 \cdot 3H_2O$ (211 g/liter)
0.5M KI (83 g/liter)
0.5M KSCN (49 g/liter)
0.2M $AgNO_3$ (34 g/liter)
Na_2CO_3, saturated
10% thioacetamide, CH_3CSNH_2 (100 g/liter) NOTE: This does not keep well. Make in small quantities.

Test Solutions for "Unknowns":
Generally speaking, these solutions should be about 0.2M or 0.1M, except for costly solutes, or those with limited solubility. Solutions should be made up very carefully. We recommend that each student be given about 15 cc of test solution in an 18 x 150 mm test tube. If he or she runs out before completing the identification of both the cation and anion, no more test solution should be supplied. Warn students about this before they begin analysis.

Some solutions on this list are not otherwise used in this course. If the substances are not already on your stockroom shelf, they may be omitted.

0.1M $AgNO_3$ (17 g/liter)
0.1M $Pb(NO_3)_2$ (33 g/liter)
0.1M $Hg_2(NO_3)_2 \cdot 2H_2O$ (56 g/liter)
0.2M $CuCl_2 \cdot 2H_2O$ (34 g/liter)
0.2M $CuBr_2$ (44.6 g/liter)
0.2M $CuSO_4 \cdot 5H_2O$ (49.8 g/liter)
0.2M $Cu(NO_3)_2 \cdot 3H_2O$ (48.3 g/liter)
0.2M $HgCl_2$ (54 g/liter)
0.1M $Hg(NO_3)_2 \cdot H_2O$ (34.2 g + 5 cc 6M HNO_3; warm and add water to make 1 liter)
0.1M $SbCl_3$ (22.8 g + 250 cc 12M HCl/liter)
0.2M $BiCl_3$ (63 g/liter)
0.2M $AlCl_3 \cdot 6H_2O$ (48 g/liter)

0.1M $Al_2(SO_4)_3 \cdot 18H_2O$ (66.6 g/liter)
0.2M $CrCl_3 \cdot 6H_2O$ (53 g/liter)
0.2M $Cr(NO_3)_3 \cdot 9H_2O$ (80.0 g/liter)
0.2M $Cr_2(SO_4)_3$ (70 g/liter)*
0.2M $FeCl_3 \cdot 6H_2O$ (54 g/liter)
0.2M $CoCl_2 \cdot 6H_2O$ (48 g/liter)
0.2M $Co(NO_3)_2 \cdot 6H_2O$ (58.2 g/liter)
0.2M $NiCl_2 \cdot 6H_2O$ (48 g/liter)
0.2M $NiSO_4 \cdot 6H_2O$ (52.5 g/liter)
0.2M $Ni(NO_3)_2 \cdot 6H_2O$ (58.2 g/liter)
0.2M $MnCl_2 \cdot 4H_2O$ (40 g/liter)
0.2M $Mn(NO_3)_2 \cdot 6H_2O$ (58 g/liter)
0.2M $ZnCl_2$ (27 g/liter)
0.2M $Zn(NO_3)_2 \cdot 6H_2O$ (60 g/liter)
0.2M $BaCl_2 \cdot 2H_2O$ (48 g/liter)
0.2M $Ba(NO_3)_2$ (52.3 g/liter)
0.2M $CaCl_2 \cdot 2H_2O$ (29 g/liter)
0.2M $SrCl_2 \cdot 6H_2O$ (53 g/liter)
0.2M $Sr(NO_3)_2$ (42 g/liter)
0.2M $NaCl$ (12 g/liter)
0.2M NaI (30 g/liter)
0.2M $NaBr$ (20.6 g/liter)
0.2M KCl (15 g/liter)
0.2M KBr (23.8 g/liter)
0.2M NH_4Cl (10.6 g/liter)
0.1M $Na_2CO_3 \cdot 10H_2O$ (28.6 g/liter)
0.1M $Na_3PO_4 \cdot 12H_2O$ (38 g/liter)
0.1M $NaNO_2$ (6.9 g/liter)
0.1M Na_2S (7.8 g/liter)
0.1M $Na_2SO_3 \cdot 7H_2O$ (25 g/liter)
0.1M K_2CrO_4 (19.4 g/liter)
0.1M $NaC_2H_3O_2$ (8.2 g/liter)
0.1M $NaClO_3$ (10.6 g/liter)

*Note: As usually received, this reagent contains a variable amount of water. Use 70 g of the anhydrous material, if possible, or estimate the amount of water.

Suggestions:

1. Each student should be encouraged to do as many "unknowns" as possible in the time allotted to this study.
2. In a period of two weeks (working one hour per day) students average about 4 or 5 unknowns.
3. It is recommended that four unknowns be required; then give extra credit for each one done above the minimum.
4. Try to give each student a variety of ions from as many groups as possible with which to work.
5. Since the cations in Group VI do not have to be removed for anion analysis, it is suggested that cations in Group VI be reserved for students who get behind and need to do a less time-consuming analysis. This is not to say that cations in Group VI should not be given to all students, but it can be a great help for the self-confidence of those who need this type of assistance.
6. Stress the directions in the laboratory manual for keeping reagents uncontaminated and for keeping good records.

Evaluation:

Each student can be evaluated on the results of his findings in the laboratory, on his ability to keep good records, upon his skill in writing equations to represent reactions, and upon his reports. Some teachers like to let the fourth (or last) unknown count as a test grade. Whether or not a written test is given over the procedure and the principles underlying qualitative analysis is the decision of each teacher.

We consider it more important that students enjoy this study and leave the course with a high level of interest and enthusiasm than that they be forced to try to work beyond their own capabilities.

APPENDICES

APPENDIX: MAKING MOLECULAR MODELS

Models of hydrogen, oxygen, water, and ammonia molecules can be made from inexpensive materials. These models are quite useful in:

a. Showing conservation of matter. If one "molecule" of oxygen and one "molecule" of hydrogen are disassembled to make two "molecules" of water, it is readily apparent that more hydrogen is required. This presents visual evidence that two molecules of hydrogen and one molecule of oxygen produce two molecules of water.

b. Showing the approximate bond angles. It is easy to see that both hydrogen and oxygen molecules are linear, that the bond angle in water is about 105°, and that the angles in ammonia are about 120° when the models are used.

c. Illustrating acid-base theory. When one begins with a water molecule and an ammonia molecule, it is easy to demonstrate the concept of proton transfer using models. Such visual assistance helps greatly in presenting these ideas.

Materials:

3 Styrofoam spheres, approximately 5.1 cm in diameter
3 Styrofoam spheres, approximately 6.4 cm in diameter
1 knife or hacksaw blade (steak knives with serrated edges are ideal for cutting Styrofoam)
8 chenille stems (commonly known as pipe cleaners)
glue
latex paint: white, red, blue
1 drawing compass (for circles and arcs)
1 metric ruler
1 protractor
1 piece of fine sandpaper

Note: Suppliers of Styrofoam spheres may be located by looking in the yellow pages of the phone book under Plastics — Foam. If there are no local suppliers, you can order from one of the following manufacturers:

Star Band Co.
Broad and Commerce Sts.
Portsmouth, Va. 23707

Plasteel Corp.
26970 Princeton
Inkster, Michigan 48141

Background:

When molecules are packed together in a solid crystal, the distance between the nuclei of two neighboring atoms on *different* molecules is one measure of the size of the atoms. For the purpose of making a simple model, we can think of the two atoms as hard spheres touching each other. For each atom, the radius of the sphere is called the *van der Waals radius*, r_v.

Within a molecule, the nuclei of two neighboring atoms are closer together than the sum of the van der Waals radii. In our model we can think of the spheres with pieces sliced off, so the nuclei can be closer together. The distance from the nucleus to the surface of the "slice" is called the *covalent radius*, r_c. The value of the covalent radius depends both on the type of atom and on the type of molecule it is in.

The following table gives values of the van der Waals radii and covalent radii of hydrogen, oxygen, and nitrogen atoms in molecules of hydrogen, oxygen, water, and ammonia.

Element	*Van der Waals radius, in meters*	*Covalent radius, in meters*			
		in H_2	in O_2	in H_2O	in NH_3
hydrogen	1.2×10^{-10}	0.37×10^{-10}		0.30×10^{-10}	0.30×10^{-10}
oxygen	1.4×10^{-10}		0.60×10^{-10}	0.66×10^{-10}	
nitrogen	1.5×10^{-10}				0.70×10^{-10}

Styrofoam spheres are available commercially from a number of suppliers. They come in fixed sizes. We recommend using the 5.1-cm spheres for hydrogen and the 6.4-cm spheres for both oxygen and nitrogen, even though the scale will be slightly off. We also recommend compromising on the covalent radii and using the same value in all compounds of a given atom, so the molecules may be disassembled to form new molecules. The following values may be used for the scale models.

Element	*Van der Waals radius*	*Covalent radius*
hydrogen	2.55 cm	0.66 cm
oxygen	3.20 cm	1.40 cm
nitrogen	3.20 cm	1.50 cm

Procedure:

Hydrogen atoms

1. Draw a circle with a radius of 2.55 cm on a piece

of paper. This corresponds to the van der Waals radius, r_v, for a hydrogen atom.

2. Draw a diameter through the center of the circle.

3. On the diameter, mark a point that is one covalent radius, r_c, from the center. Through this point construct a perpendicular that intersects the circle (see Diagram 1).

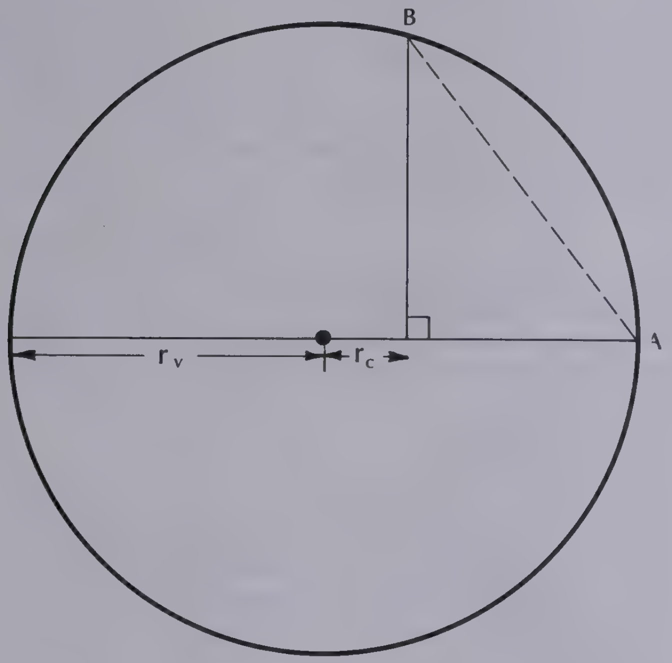


Diagram 1 (not to scale)

4. Set a compass for the distance AB on the paper circle (see Diagram 1) and draw a circle on one of the 5.1-cm spheres, using this radius. (Note: As a base for the compass use a small piece of thin cardboard held in place by transparent tape. Otherwise, the point of the compass will penetrate the Styrofoam and distort the circle to be drawn on the sphere. Draw identical circles on the remaining 5.1-cm spheres.

5. Slice off the marked portions of each sphere, allowing an extra 3 to 4 mm on the larger portions for smoothing. Smooth the cut surface with sandpaper. *Caution:* Work carefully, as Styrofoam is fragile. Stop when each sphere is worked down to the marked circle.

6. Paint the *curved* surface of each "hydrogen atom" with white latex paint, leaving the flat surfaces unpainted. Wood applicator sticks supported in test tube racks make excellent supports for the models as they dry.

7. When the models are dry, embed a chenille stem about halfway into the center of the flat surface of two of the "hydrogen atoms"; reserve the third for later.

Oxygen atoms

1. To shape the oxygen atoms, follow the procedure for the hydrogen atoms, but use r_v = 3.20 cm and r_c = 1.40 cm. Use two Styrofoam spheres of diameter 6.4 cm.

2. Paint the *curved* surface of each "oxygen atom" with red latex paint, leaving the flat surfaces unpainted.

3. When the models are dry, embed a chenille stem about halfway into the center of the flat surface of each "oxygen atom."

Nitrogen atom

1. Follow steps 1, 2, and 3 for the hydrogen atoms, but use r_v = 3.20 cm and r_c = 1.50 cm.

2. On the paper circle, draw a radius at an angle of 120° to the diameter. Connect the tip of this radius to the end of the diameter (see Diagram 2). This is the distance AC.

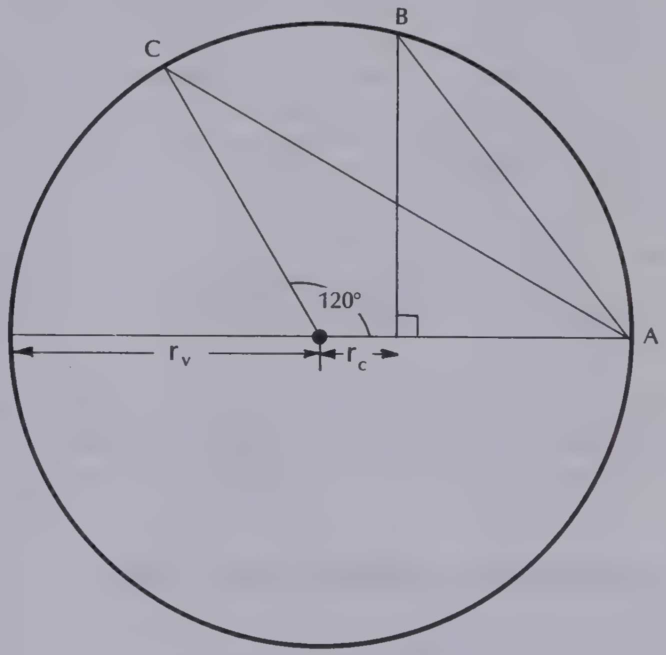


Diagram 2 (not to scale)

3. Mark a dot on the remaining 6.4-cm sphere. Set AC on the compass and lay off this distance from the first dot. Make a second dot at this point.

4. Using these two dots as centers and the distance AC as the compass setting, locate one more dot equidistant from the first two. (Note: Draw short arcs from each of the original two dots to locate the third dot.)

5. Now set the distance AB on the compass. Draw a circle around each of these three dots, using a thin cardboard support for the compass point, as before.

6. Cut off the three marked portions, leaving 3 to 4 mm for smoothing. Smooth down to the circles with sandpaper.

7. Paint the curved surface of the "nitrogen atom" with blue latex paint. Attach chenille stems to each of the flat surfaces, as before.

Assembling the hydrogen molecule

1. Place the flat surfaces of the two stemmed "hydrogen atoms" against each other.
2. Twist the chenille stems together to hold in place.

Assembling the oxygen molecule

1. Place the flat surfaces of the two "oxygen atoms" against each other.
2. Twist the chenille stems together to hold in place.

Assembling the water molecule

1. Hollow out the unpainted surface of the unstemmed "hydrogen atom" so that it will fit against the curved surface of one of the "oxygen atoms."
2. Place the flat surface of one "oxygen atom" against the flat surface of one "hydrogen atom." Twist the chenille stems together to hold in place.
3. Place the hollowed-out "hydrogen atom" along the "oxygen atom" and slide it over until it is about 1 mm from the other "hydrogen atom." This arrangement corresponds to the 105° bond angle in the water molecule.

Assembling the ammonia molecule

1. Attach a chenille stem to the center of the hollowed-out surface of the unstemmed "hydrogen atom."
2. Place the three "hydrogen atoms" against the flat surfaces of the "nitrogen atom" and twist the chenille stems to attach in place. This arrangement corresponds to the 120° bond angles in ammonia.

APPENDIX: EQUIPMENT AND SUPPLIES LIST

EQUIPMENT

A course in which students proceed at their own pace does not require as much reusable equipment as a course in which the entire class works on the same experiment at the same time. For this reason, equipment is listed with the number of items required for one student. If you need to order specialized equipment for this course, figure that not more than five to ten students will be working on the same experiment at the same time. However, basic equipment items such as test tubes, burners, and ringstands are used so frequently that, if financially possible, each student should be allowed for in figuring the total number needed. In addition, there must be an allowance for glassware breakage.

Glassware or Plasticware

Item	*Quantity per student*
Beaker, 50 cc	2
100 cc	1
250 cc	2
400 cc, tall form	6
1000 cc	1
Bottle, gas-collecting, 60 cc	1
Cover plate	1
Flask, Erlenmeyer, 250 cc	1
Funnel, 75 mm top diam. and 75 mm stem	1
Graduated cylinder, 10 cc	2
50 cc	1
100 cc	3
Jar or vial, wide-mouth, with cap, approx. 100 cc capacity (for demo. in Exp. 4-1)	24
Pipet, dropper	1
Rod, stirring	1
Slide, microscope	6
Spot plate	4
Test tube, 13 X 100 mm	18
18 X 150 mm	12
Tube, 1 meter long and 2 cm in diameter	1
Tube, gas-collecting, 50 cc, graduated	1
Vial, plastic, 4 cm by 2 cm, with cap	15
Watch glass	5

Hardware

Item	*Quantity per student*
Asbestos mat	1
Balance, centigram, ±0.01 g	1
platform	1
Barometer, mercurial	1
Blowpipe	1
Burner	1
Clamp, buret	2
Charles' Law apparatus	1
Conductivity apparatus (for demo. in Exp. 4-1)	1
Electrolysis apparatus, Hoffman type (for demo. in Exp. 8-3)	1
Evaporating dish	1
Forceps, pair	1
*Grating, diffraction, 2 cm x 2 cm	1
Magnet, bar	1 pair
horseshoe	1
Magnifying lens	1

*Purchasable in 8½" x 11" sheets which can be cut into squares and mounted between two microscope slides for use with students. Edmund Scientific Co., Barrington, N.J. 08007 is an economical source.

Item	*Quantity per student*
Metric ruler, 15 or 30 cm	1
Meter stick	1
Microscope, binocular	1
Multimeter (volt-ohmmeter) (for demo. in Exp. 4-1)	1
Petri dish	3
Ringstand	2
Spoon, deflagrating	1
Stopper, cork, any size	1

Item	*Quantity per student*
Stopper, rubber, #8	1
rubber, to fit gas-collecting tube	1
Syringe, plastic, 25 cc capacity	1
Test tube brush	1
Test tube holder	1
Test tube rack	1
Thermometer, range −10°C to 110°C	2
Tongs, beaker	1 pair
Tongs, crucible	1 pair

SUPPLIES

Chemical Supplies

The indicated quantities and standard of purity are based on what quantity and grade are both suitable and economical for establishing a stockroom supply of chemicals for LABORATORY KEYS TO CHEMISTRY. In many cases, the economic purchase will create a ten-or-more years' supply.

Item	*Purchased as*	*Quantity*
Acid, acetic	glacial, reagent	5 lb
hydrochloric	concentrated, reagent	12 lb
nitric	concentrated, 90%, reagent	7 lb
sulfuric	concentrated, reagent	18 lb
Base, ammonia water	concentrated, reagent	10 lb
calcium hydroxide	tech., powder	1 lb
potassium hydroxide	lab gr., pellet	2 lb
sodium hydroxide	lab gr., pellet	4 lb
Acetone	tech.	1 gallon
Aluminum chloride	$AlCl_3 \cdot 6H_2O$, lab. gr.	1 lb
Aluminum nitrate	$Al(NO_3)_3 \cdot 9H_2O$, reagent	1 lb
Aluminum sulfate	$Al_2(SO_4)_3 \cdot 18H_2O$, tech., granular	1 lb
Ammonium carbonate	purified powder	1 lb
Ammonium chloride	tech. gr.	1 lb
Ammonium nitrate	tech. gr.	1 lb
Ammonium oxalate	reagent gr.	1 lb
Ammonium sulfide	light solution	2 pts at least
Antimony (III) chloride	reagent	¼ lb
Barium chloride	$BaCl_2 \cdot 2H_2O$, reagent, crystal	1 lb
Barium nitrate	reagent, crystal	1 lb
Benzene	pure, industrial	1 gallon
Bismuth nitrate	$Bi(NO_3)_3 \cdot 5H_2O$, reagent, crystal	¼ lb
Buffer solution, pH 7		1 pt
Calcium metal	purified, turnings	¼ lb
Calcium acetate	purified, powder	1 lb
Calcium carbonate	marble chips	5 lb
Calcium chloride	anhydrous, 4 mesh for desiccator	1 lb
Calcium nitrate	$Ca(NO_3)_2 \cdot 4H_2O$, purified, flake	1 lb
Calcium oxide (lime)	lump	1 lb
Carbon disulfide		1 or 2 pt
Carbon tetrachloride	tech. gr.	2 gallons
Charcoal	wood, lump	5 lb
Charcoal, wood, block		3 doz/class
Chromium (III) nitrate	$Cr(NO_3)_3 \cdot 9H_2O$, reagent, crystal	1 lb

Item	*Purchased as*	*Quantity*
Cobalt (II) chloride	crystal, tech.	¼ lb
Cobalt (II) nitrate	purified, crystal	¼ lb
Copper sheet, soft, light	B & S 30, 12″ x 12″, about ½ lb per sq ft	1 sheet
Copper wire, soft, bare	16 gauge	4 oz spool
Copper wire, soft, bare	30 gauge	4 oz spool
Copper (II) chloride	purified, crystal	1 lb
Copper (II) nitrate	purified, flake	1 lb
Copper (II) sulfate	tech., crystal (hydrate)	5 lb
Copper (II) sulfate	anhydrous powder	¼ lb
Ethanol	ethyl alcohol, 95% denatured	2 gallons
Filter paper, 11 cm dia.	suitable for qual. analysis, 100 sheets/pkg	25 pkgs
Iodine	tech. gr.	1 lb
Iron filings, fine		2 lb
Iron (III) chloride	$FeCl_3 \cdot 6H_2O$, reagent	1 lb
Iron (II) sulfate	$FeSO_4 \cdot 7H_2O$, reagent,	1 lb
Lead (II) nitrate	tech., crystal	2 lb
Litmus paper, neutral		12 vials
Magnesium metal	ribbon, purified	2 oz
Magnesium nitrate	$Mg(NO_3)_2 \cdot 6H_2O$, reagent, crystal	1 lb
Magnesium sulfate	$MgSO_4 \cdot 7H_2O$, lab gr., crystal	1 lb
Manganese dioxide	tech., powder	1 lb
Manganese (II) nitrate	$Mn(NO_3)_2 \cdot 4H_2O$, reagent	1 lb
Mercury metal	tech.	1 lb
Mercury (II) chloride	reagent, crystal	¼ lb
Mercury (I) nitrate	$Hg_2(NO_3)_2 \cdot 2H_2O$, reagent, crystal	1 oz
Mercury (II) oxide, red	purified	1 oz
Methanol	reagent	1 gallon
Methyl orange	reagent, powder	1 oz
Nichrome wire		3 to 5 ft
Nickel (II) chloride	tech.	1 lb
Nickel (II) nitrate	purified, flake	1 lb
Nickel (II) sulfate	tech.	1 lb
pH test paper	wide range (Hydrion)	
	Type A	12 vials
	Type B	12 vials
Phenolphthalein	powder, lab gr.	¼ lb
Potassium chloride	lab gr.	1 lb
Potassium chromate	reagent, crystal	1 lb
Potassium ferrocyanide	tech.	1 lb
Potassium iodate	reagent, granular	¼ lb
Potassium iodide	lab gr.	1 lb
Potassium nitrate	tech.	2 lb
Potassium permanganate	lab gr., crystal	1 lb
Potassium thiocyanate	reagent, crystal	¼ lb
Silver nitrate	reagent, crystal	1 lb
Splints	thin wood	2 pkg of 500
Sodium acetate	tech., anhydrous	1 lb
Sodium bicarbonate	tech., powder	1 lb
Sodium bisulfite	tech., granular	1 lb
Sodium borate (borax)	tech.	1 lb
Sodium bromide	reagent, crystal	1 lb
Sodium carbonate	tech., monohydrate	2 lb
Sodium chlorate	tech.	1 lb
Sodium chloride	lab gr., crystal	5 lb
Sodium chromate	tech. anhydrous	1 lb
Sodium iodide	lab gr., crystal	¼ lb

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